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CHEMISTRY 30

MODULE 4: THE PRACTICAL SIDE OF REDOX

Distance Learning

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Chemistry 30

Module 4

The Practical Side of Redox

Distance Learning

Alberta EDUCATION

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Chemistry 30
Student Module
Module 4
The Practical Side of Redox
Alberta Distance Learning Centre
ISBN 0-7741-0946-7

Cover Photo: WESTFILE INC.

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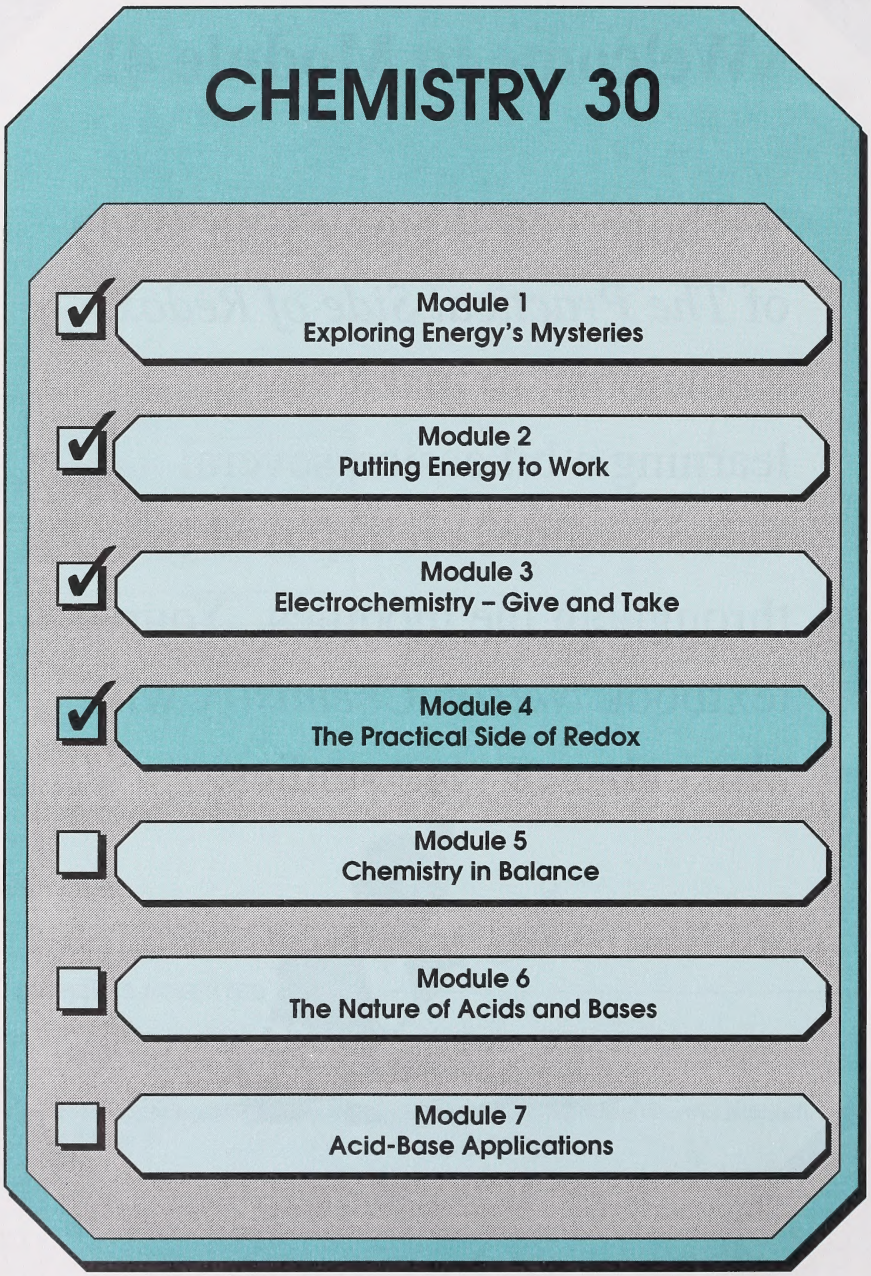
Welcome to Module 4!

We hope you'll enjoy your study of *The Practical Side of Redox*. Remember, to make your learning a bit easier, several videocassettes are referred to throughout the modules. Your textbook *Nelson Chemistry* will also enhance your studies.

COURSE OVERVIEW

This course contains seven modules. The module you are working in is highlighted in deeper colour. It is recommended that you work through the modules in the order given since several concepts build on each other as you progress in the course.



MODULE 4

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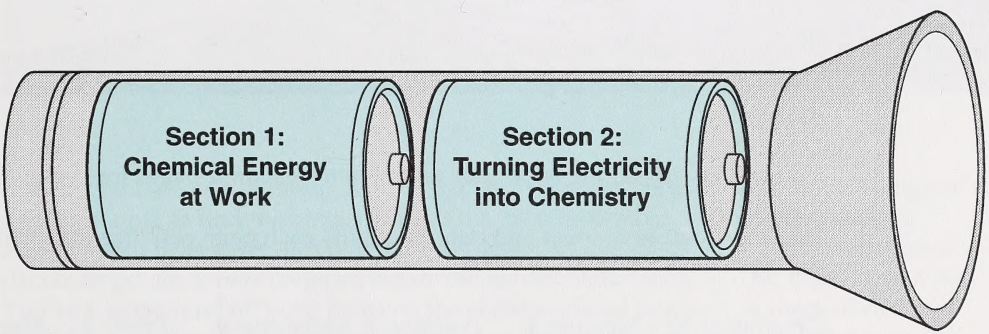
MODULE OVERVIEW

Have you ever bought a battery? And have you ever spent a loonie? At first glance, you might say that the only relationship between a battery and a loonie is that you can pay for batteries with loonies. But from a chemical point of view, these two items are remarkably similar because of the electrochemical processes that are involved with each of them.

How many things can you think of that run on batteries? Society depends heavily on this technological application of redox for everything from operating toys to computers to cellular phones. But how do batteries actually work? Is it a complicated process?

In Module 3, you learned about electrochemical reactions and their connection with processes such as metallurgy (for example, producing loonies). In the first part of this module, you will learn about some practical electrochemical processes including those that occur in batteries – the way they work and the way they are constructed. In the second half of this module, you will learn about the process and applications of electrolysis, as well as perform calculations relating to various electrochemical processes.

Module 4: The Practical Side of Redox



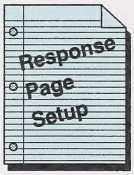
Evaluation

Your mark in this module will be determined by how well you complete the assignments at the end of each section. You must complete all assignments. In this module you are expected to complete two section assignments and one final module assignment. The mark distribution is as follows:

Section 1 Assignment	**40 marks**
Section 2 Assignment	**40 marks**
Final Module Assignment	**20 marks**
TOTAL	**100 marks**

When doing your assignments, work slowly and carefully. If you are having difficulties, go back and review the appropriate section.

Read all parts of your assignment carefully. Plan and do your rough work on your own paper. Revise and edit your responses; then set up your final copy for submission on your own paper. Lined looseleaf is recommended. Make sure your answers are neat and organized, with wide left margins and space for teacher comments after each assignment.



When you see this icon, ideas and details are provided to help you set up and organize your answer in a certain way.

Before submitting your responses, be sure to proofread them carefully to ensure they say what you want, that they are neat and clear, and that they are complete and missing no material.

You will be submitting **only** your **assignment response pages** for evaluation.

It is important to number and clearly identify each page with this information placed at the top.

Chemistry 30 – Module 4 Section # Assignment Page # Name and ID #

SECTION 1

Chemical Energy at Work

Imagine you are on a camping trip. It is dusk. You take off your walkman and dig out a flashlight only to find the batteries are dead. You remember having seen a service station about fifteen minutes down the highway, so you check your digital watch and decide to get some new batteries before the service station closes. You walk to your car. You kick some mud off your boots on the chrome-plated bumper. A piece of rust falls from the wheel well. The engine groans when you try to start your car, but it runs on the second try. At the service station, you have to decide what type of battery to purchase – carbon? alkaline? rechargeable? AAA? AA? C? D? mercury? mercury-free? You make your choice and pay the attendant with your last few loonies. You then return to your campsite.

Could you count at least eight ways in which electrochemical reactions may be involved in your daily life? In this section, you will learn the basic principles involved in the operation of different types of batteries, as well as investigate how the voltage of a battery is determined. You will also examine and evaluate ways electrochemical reactions are used to benefit society.

ACTIVITY 1

Cells and Batteries

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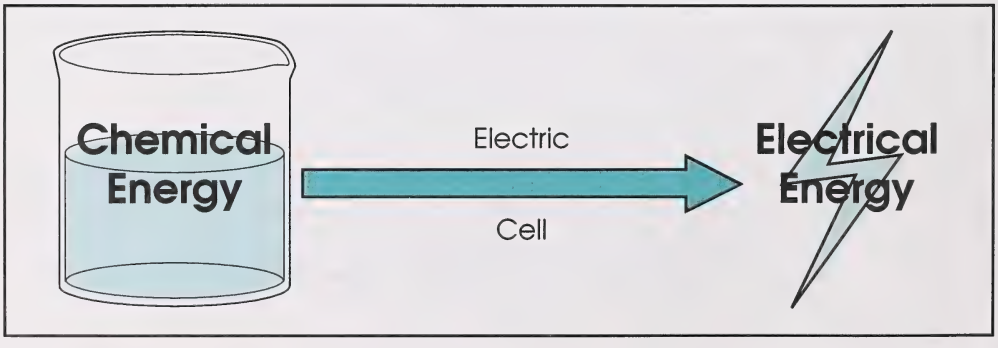
Have you ever chewed on a piece of aluminum foil? If you have metallic-based fillings, you no doubt got a shock. Why? This is just one example of change occurring in a system in which electrons are transferred from one substance to another.

In Module 3, you became familiar with the process of electron transfer in redox reactions. This module will focus on how you can put those electrons to work by converting chemical energy to electrical energy.



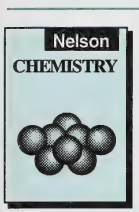
multimeter – a device that can be used to measure a variety of electrical quantities such as current, voltage, and resistance

To see if you are making those electrons work, you need a device to measure the flow of electrons. One such device is a **multimeter**. You will be using the multimeter often as you work through this module, so it is important that you know how to use it. Work through the following short investigation.

Investigation: Using a Multimeter

To learn how to use the multimeter to measure current, read the section Using a Multimeter: Conductivity Measurements of Solutions found in Appendix C on pages 531-532 of your textbook. Further guidelines follow. Pay special attention to the required components, safety aspects, and applied science skills.

Nelson CHEMISTRY



Science Skills

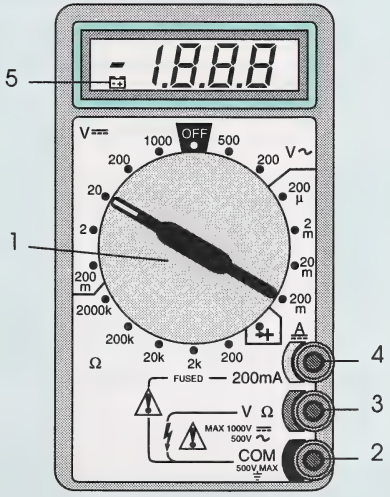
- ☐ A. Initiating
- ☑ B. Collecting
- ☐ C. Organizing
- ☐ D. Analysing
- ☐ E. Synthesizing
- ☐ F. Evaluating

Guidelines

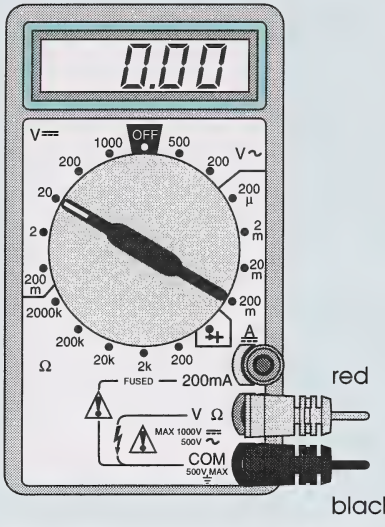
- Follow the directions using a cup of tap water as your control and a cup of salt-water as your test solution.
- Make the salt solution by dissolving 5 mL (one level teaspoon) of salt in 200 mL of tap water.
- General instructions for using the multimeter are given as follows:

Parts of the Multimeter

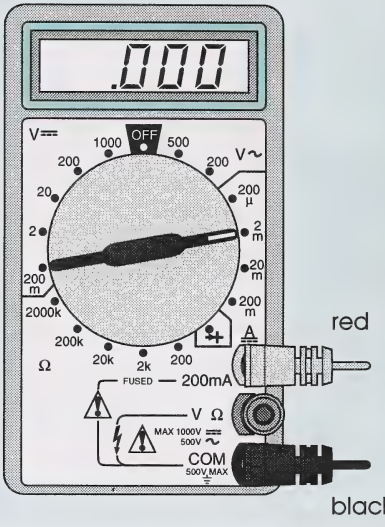
1. Function/Range switch – This switch is used to select direct current volts (V⎓), direct current amperes (A⎓), or resistance ohms (Ω).
2. Black lead terminal – Your black lead should always be in this terminal.
3. Red lead terminal – Your red lead should be in this terminal when measuring **voltage (potential difference – V⎓)**.
4. Red lead terminal – Your red lead should be in this terminal when measuring **current (amperes – A⎓)**.
5. Low battery indicator – This indicates when your battery needs changing.



Proper Connections for Measuring Voltage less than 20 V (V⎓)



Proper Connections for Measuring Current less than 2 mA (A⎓)



- You can measure the relative conductivity of solutions either by resistance, Ω (poor conductor → high resistance; good conductor → low resistance); or by current, A⎓ (poor conductor → low current; good conductor → high current), if a power source is attached.

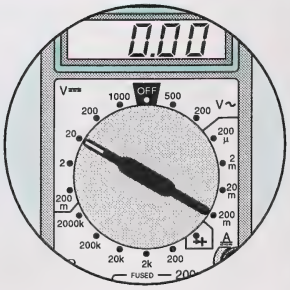
Caution

- Make sure you turn the multimeter off when not in use.

1. Record the results of your conductivity test for the water and salt solution.

Check your answers by turning to the Appendix, Section 1: Activity 1.

Measuring voltage using a multimeter is similar to the preceding procedure, except the dial is set to the direct current volts (DCV) scale (V⎓) as shown in the preceding diagram and the diagram at the side. In addition, if the needle moves left instead of right, you will need to reverse the connections you have made to your **electrodes**.



electrodes – solid material through which electricity flows; usual site of electrical connections

To become more familiar with the use of a multimeter, practise using it as you design a simple electric cell in the following investigation.

Investigation 13.1: Demonstration of a Simple Electric Cell

Nelson CHEMISTRY

Carefully read Investigation 13.1: Demonstration of a Simple Electric Cell found on page 387 of *Nelson Chemistry*. Pay special attention to the required components, safety aspects, and applied science skills.

Science Skills

- ☐ A. Initiating
- ☑ B. Collecting
- ☐ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☑ F. Evaluating

Guidelines

- Before using the metals, polish them with steel wool to remove the oxide coating.
- Make sure that the two metals do not come in contact with each other, or you will not get proper results.
- If a silver strip is not available, use an iron nail instead.
- Pay special attention to Figure 13.2 at the top of page 388 of your textbook to see how the multimeter should be connected to the metal strips. The multimeter will be used instead of a voltmeter and ammeter.

Nelson CHEMISTRY

- To measure the voltage, follow the directions under the subheading Voltage Measurements of Batteries found on page 532 of your textbook.
- For measuring voltage, set your multimeter to V ⎓ , 20 V range; for measuring current, set your multimeter to A ⎓ , 20 mA range.

2. Record your results in your notebook in a data table similar to the following. You should have a minimum of ten results in your data table.

Trial	Metal 1	Metal 2	Material Between Metals	Ammeter Reading (milliamperes)	Voltmeter Reading (Volts)
1	copper	zinc	salt water/towel		

voltmeter – a device used to measure the potential difference between any two points in a circuit

ammeter – a device used to measure electrical current

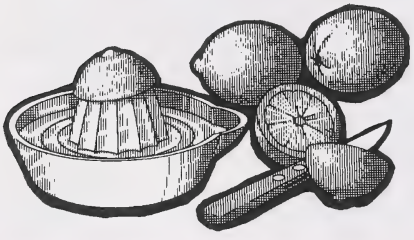
batteries – two or more voltaic cells connected in series, or one or more voltaic cells in a finished package

3. What evidence was there that a chemical reaction was occurring?

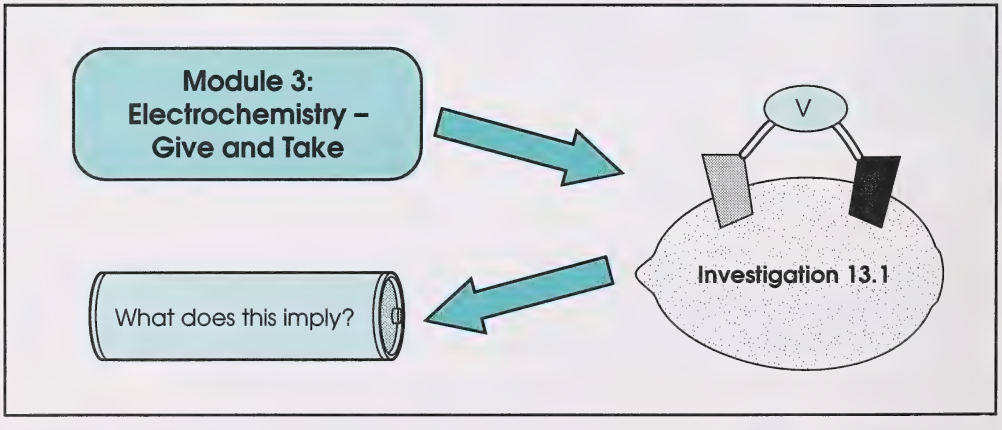
4. Evaluate Luigi Galvani's prediction. To what extent was it correct and to what extent was it incorrect?

5. In an electrochemical reaction, electrons are transferred directly by contact. Why was it necessary to keep the metals from touching and to connect them with a wire (to which the voltmeter or ammeter or multimeter was attached)?

6. If you had substituted a nonconducting material such as a block of wood for the salt-soaked towel or the fruit, the cell would not have worked. Why was it necessary that the material between the two metals be conductive?



Check your answers by turning to the Appendix, Section 1: Activity 1.



electrolyte – a substance that, when dissolved in water, conducts an electric current

In the preceding investigation, you learned that one major difference between the electrochemical reactions studied in Module 3 and the electrochemical reactions you are studying in this module is that in this module, the reactants are separated. Therefore, the electron transfer that would normally occur in a chemical change by contact is now forced to occur through a wire. An electrolyte then becomes necessary to complete the circuit. When this happens, the moving electrons which have energy are able to do work. This is illustrated in the following diagram.

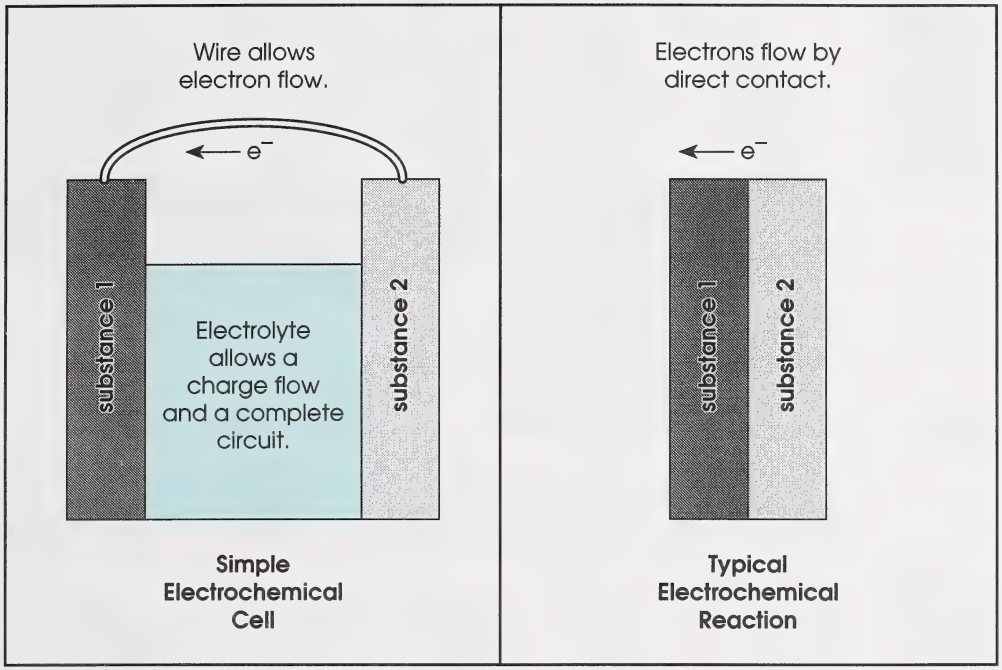


FIGURE 4.1: A Simple Cell Compared to an Electrochemical Reaction

Nelson CHEMISTRY

But what is the major difference between the simple **electrochemical cell** you made in Investigation 13.1, and a 9 V battery used in such items as smoke detectors? To find out, read all of page 388 in your textbook and answer the next two questions in your notebook.

electrochemical cell – any electrochemical reaction in which an external current is made from a chemical reaction or a chemical reaction is made from an electric current

7. a. Of the following five terms, which three usually mean the same thing?

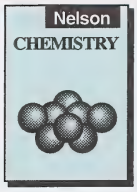
   - electric cell
   - electrochemical cell
   - battery
   - voltaic cell
   - galvanic cell

   b. Which of the following would be analogous to a battery? Which would be analogous to a voltaic cell?

   - a string of pearls
   - a stack of dimes
   - a doughnut
   - a traffic jam

8. You have a box of copper strips and a box of lead strips. In addition, you have a roll of copper wire and the following electrolyte materials:

Electrolyte Material	Number Available	Cell Voltage Produced
lemon	21	0.86 V
apple	60	0.20 V
potato	60	0.15 V



Using as few materials as possible, construct individual cells from the materials described. Hook the cells in **series** as shown in Figure 13.4 on page 388 of your textbook to make batteries that could generate the minimum voltage required to operate each of the devices shown in the following table. In each case, use one type of cell only (lemons **or** apples **or** potatoes, but not mixed).

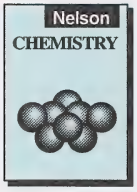
series – the positive electrode of one cell hooked to the negative electrode of another cell so that the cell voltages are additive

Object to Operate	Voltage Required	Amount and Type of Electrolyte Required	Voltage Generated
car battery	12 V		
camera flash	6 V		
transistor radio	9 V		
flashlight	3 V		
ghetto-blaster	9 V		

Check your answers by turning to the Appendix, Section 1: Activity 1.

voltaic cells – any cells that produce an electric current from a spontaneous electrochemical reaction

From the previous question, it is obvious that modern batteries and **voltaic cells** are much more efficient than those made out of household materials. As the knowledge about cells and batteries has increased, so has the terminology used to discuss them. Therefore, it is necessary to examine some of the terminology used when talking about cells and batteries.



Read page 389 from *Nelson Chemistry*, paying particular attention to Table 13.1 at the bottom of the page; then, look carefully at Figure 13.6 on page 390 to help you understand the terminology of cells.

Note: The term battery is often incorrectly used to describe a 1.5 V dry cell, such as AA cells or D cells.

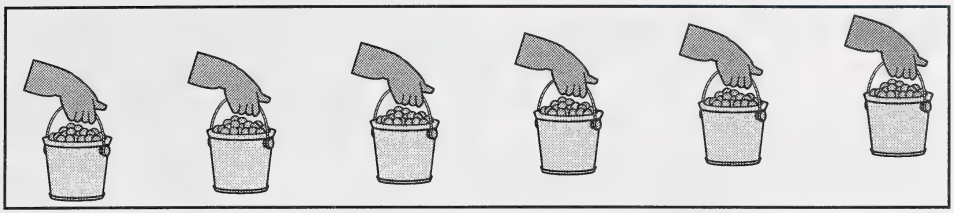
To see if you understand the material you have looked at in this activity, complete the following questions.

9. Answer textbook questions 1, 2, 4, and 5 on page 390 of *Nelson Chemistry*.

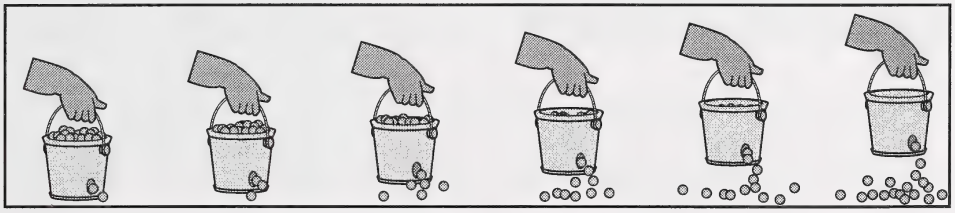
10. The following analogy, using a bucket and marbles, is a simple system that relates to a voltaic cell.

    Fill in the blanks of the given analogy using the following terms (terms may be used more than once):

    - anode
    - potential difference
    - charge
    - remain the same
    - cathode
    - power
    - total energy
    - current
    - energy density
    - increase



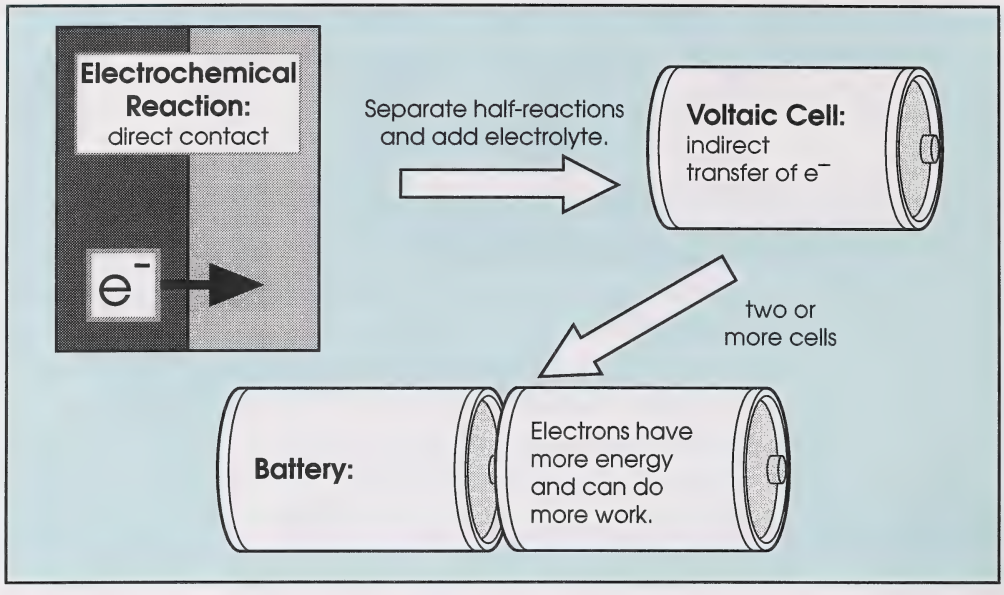
A bucket is filled with marbles which represent electrons. As the bucket is raised off the floor, each marble gains potential energy. The higher the bucket is raised, the greater the ____a.____ of the marbles becomes. However, the total ____b.____ in the bucket remains constant because the number of marbles does not change. If more marbles were added to the bucket, the charge would ____c.____, and the potential difference would ____d.____ because potential difference depends on the energy of each marble, not the total potential energy of all the marbles in the bucket.



If a hole were made in the bucket, the marbles would fall to the floor. The rate at which the marbles fall to the floor would represent the ____e.____. The bucket would then be considered to be the ____f.____ of this model voltaic cell while the floor would be considered to be the ____g.____ because electrons always flow from the ____h.____ to the ____i.____. The ____j.____ that this model cell would deliver would be based on the height the marbles are above the floor and the rate at which the marbles fall, which is why the amount of power used is determined by the rate at which the energy is being transferred to the floor. If the mass of each marble decreases and its potential energy remains the same, then the ____k.____ of the marbles increases which means that a bucket filled with the same mass of marbles could last longer because there is more ____l.____ and ____m.____ available to be transferred.

11. Explain why you experience a shock if you have metallic-based fillings and you chew on a piece of aluminum foil.

Check your answers by turning to the Appendix, Section 1: Activity 1.



In this activity, you learned about the basic components of a voltaic cell, the difference between a cell and a battery, and some of the terminology associated with electric cells. By constructing simple cells from household materials, you saw that the main difference between a voltaic cell and the electrochemical reactions discussed in Module 3 is that a voltaic cell is a system in which the half-reactions are separated so that the electron transfer must occur through an external circuit (a wire) as the chemical change occurs. As the electrons are transferred through the wire, they can do work. The work done by electrons moving in the wire is what makes your flashlight turn on or your car start.

ACTIVITY 2

Technology and Redox Cells

If you had to choose between an alkaline cell and a rechargeable cell to replace the cells in your calculator, which would you choose?

Obviously, each type of cell or battery available in the marketplace today has its strong points and its weak points. Deciding on the cell or battery that's best for the job can become quite difficult, especially as technology continues to improve battery designs.

NASA

In this activity, you will improve on the simple voltaic cell you constructed in Activity 1 in order to become more familiar with some of the methods used to improve voltaic cell designs of commercially available cells. You will also compare different commercial cells to evaluate the advantages and disadvantages of each.

Nelson CHEMISTRY

The general procedure followed when improving on a technological design is often one of trial-and-error based on educated guesses. This procedure is discussed on page 390 and in Appendix C on page 528 of *Nelson Chemistry*. Read these sections on technological problem solving; then answer the following question in your notebook.

1. What is the best way to be systematic when trying to improve upon an experimental design?

Check your answers by turning to the Appendix, Section 1: Activity 2.

Nelson CHEMISTRY

Now that you have read about how to improve technology using the problem-solving model discussed in your textbook, try this procedure by performing Investigation 13.2: Designing an Electric Cell on pages 390-391 of your textbook.

Investigation 13.2: Designing an Electric Cell

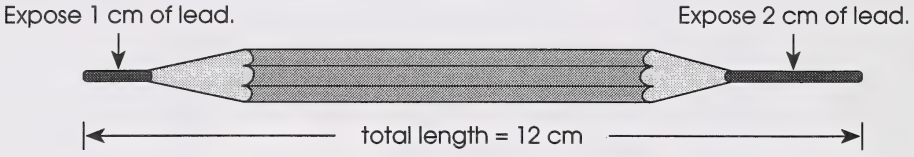
Science Skills
- ☑ A. Initiating
- ☑ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☑ F. Evaluating

Carefully follow the directions given for this investigation. Pay special attention to the required components, safety aspects, and applied science skills.

Guidelines

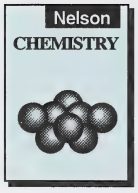
- In your procedure, make sure you remove the oxide coating from the metals with steel wool so you will get proper results.
- The top of your aluminum can must be removed. This may be done with a can-opener or tin snips.

- Use steel wool to remove the coating from the inside of the can. Be careful of the cut edges. You might try holding the steel wool with tongs to avoid injury.
- Make sure you carefully check all connections before you take a voltage reading.
- Electrodes can be made from wooden pencils. If you use a wooden pencil, you will need to make sure the graphite core of the pencil lead is exposed at both ends of the pencil as shown in the following diagram.

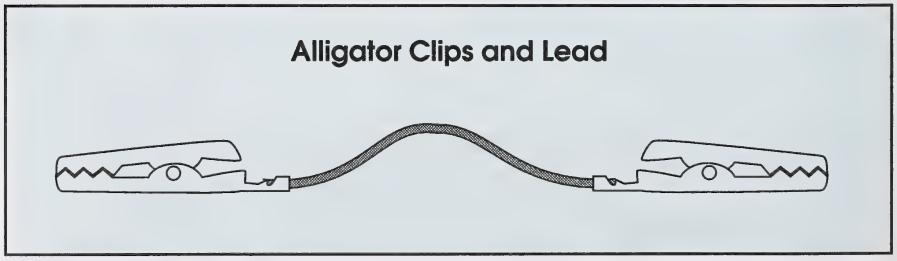


- The following table provides a list of materials you may wish to use in your investigation.

Electrolytes	Electrodes
• salt solution (Dissolve 5 mL or 1 tsp. salt in 150 mL of water.) • vinegar (Use 150 mL.) • baking soda solution (Dissolve 5 mL or 1 tsp. baking soda in 150 mL of water.)	• copper wire • lead fishing weight • medium to thick pencil lead • galvanized metal • iron nail



- Hook up your multimeter as shown in the margin on page 391 of your textbook. Use alligator clips and leads to hook the electrodes of the aluminum-can cell to the leads from the multimeter.



2. From your previous investigation and other chemistry you have learned, make a prediction as to the cell design you think would provide the greatest potential difference.

3. List the steps you will follow in the procedure.

Check your answers by turning to the Appendix, Section 1: Activity 2.

4. Carry out the procedure you developed in the previous question, recording your evidence in an appropriate table in your lab notebook.

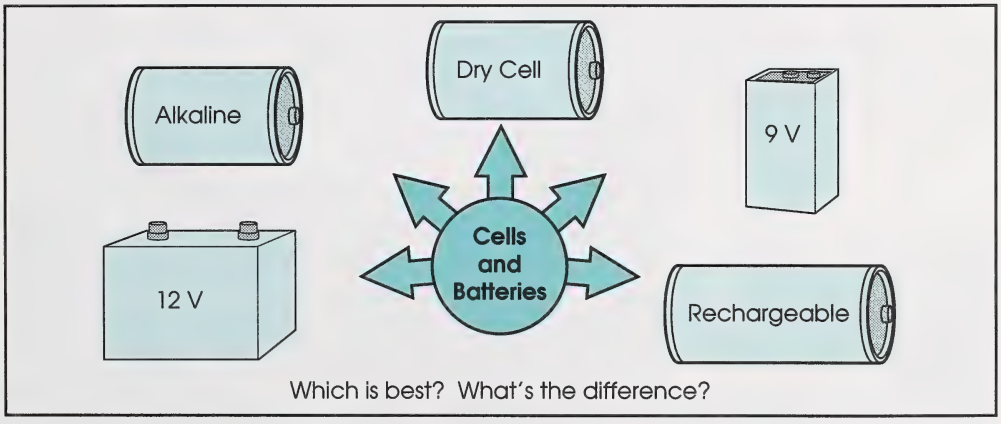
5. a. Based on the results obtained, which electrolyte and electrode would be best for your aluminum-can cell?

   b. Taking into account cost, availability of materials, and other factors, which electrolyte and electrode would be best for your aluminum-can cell?

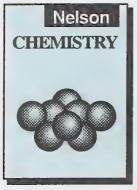
   c. How do the influencing factors determine what technological developments will be pursued?

6. Based on your results, propose at least two modifications to your aluminum-can cell that may improve its design.

Check your answers by turning to the Appendix, Section 1: Activity 2.



You may have noticed that the voltages you obtained in your aluminum-can cell were small, although some of them may have been close to or slightly greater than the 1.5 V generated by many commercial cells. Your big car battery delivers 12 V of electric potential and some small batteries deliver 9 V of electric potential. How do they do this when a 9 V battery is similar in size to a 1.5 V cell? Not only that, but what's the real difference between a 1.5 V rechargeable cell, a 1.5 V dry cell, and a 1.5 V alkaline cell?



To find the answer to these and other questions, read pages 391-393 of your textbook, paying special attention to Table 13.2 on page 392 and stopping at the Exercise on page 393. There is obviously much more information about various cells that could be discussed other than what you have just read, but some main principles were presented. Answer the following questions to help you analyse the information more closely.

7. For each of the following clues, decide whether the phrase best describes a primary cell, a secondary cell, or a fuel cell.

   a. cannot be recharged

   b. needs a continuous supply of material to operate

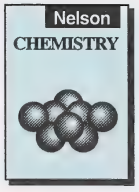
   c. several nonreversible side reactions during cell discharge

   d. discharges more quickly than primary cells doing the same job

   e. has a very high energy density

   f. most inexpensive battery to produce

8. Omar tests some batteries by placing NO-DIE alkaline cells into one cassette recorder and SURE-CHARGE carbon cells into another cassette recorder. Very soon, the SURE-CHARGE cells die. Omar concludes that the NO-DIE brand of cell is better. Stacey claims that SURE-CHARGE cells always last longer. Explain how this could be possible.

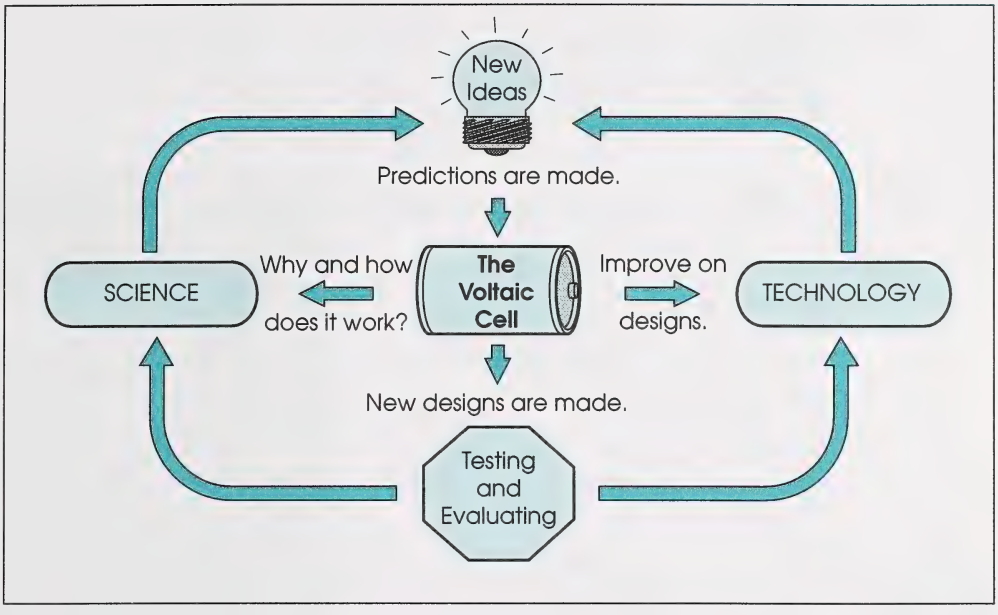


9. Answer textbook questions 6, 10, and 11 on page 393 of *Nelson Chemistry*.

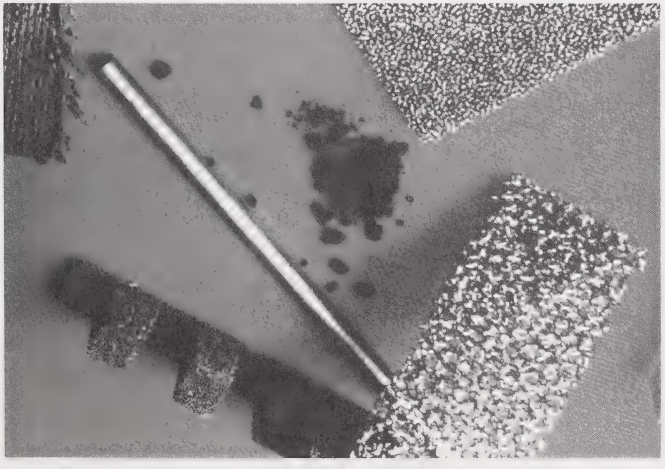
10. Would it be better to buy an alkaline cell or a rechargeable cell to replace the cells in your calculator? Why?

Check your answers by turning to the Appendix, Section 1: Activity 2.

From what you have read, it is obvious that there are many different cells and batteries available to the consumer. As new scientific principles are discovered, technology is advanced. Likewise, as new technology is developed, new scientific principles are discovered.



In this activity, you learned how to improve the technological design of the simple voltaic cell you designed in Activity 1. In the process, you realized that one difference between technological problem solving and scientific problem solving is that, in technological problem solving, you are only concerned with improving the efficiency of a device. In scientific problem solving, you are predominantly concerned with asking *how* and *why* questions. This helps explain why so many different commercially available cells and batteries exist and why more are being designed. You evaluated some of the various designs of cells and batteries and made decisions as to which type of voltaic cell was best for particular purposes.



NASA

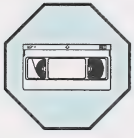
Scientists are continually working on improving designs. One recent improvement by NASA involves producing special carbon/graphite materials (see photograph) which offer better efficiency and extended lifetimes for batteries, fuel cells, and electrodes.

ACTIVITY 3

The Magic of Voltaic Cells

Imagine a cell or battery that never died. It's a consumer's dream. Technology and science have worked for almost 200 years at extending the life of cells and batteries. The result is the wide range of cells and batteries you have today. So exactly how do cells and batteries work? What happens to the electrodes? And why do the electrons move through the external circuit (wire) and not through the solution?

DAVE MORRIS



For answers to some of the preceding questions, and an introduction to batteries, view video segment 1, The Building Blocks of Electrochemistry, from the series *Electrochemistry* (TVO/ACCESS Network). Check your understanding by answering the following questions.

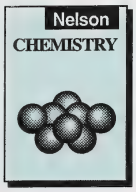
1. What name is given to the negative terminal (electrode) at which oxidation occurs?

2. What name is given to the positive terminal (electrode) at which reduction occurs?

3. What are cations? anions? Which electrodes do they migrate to?

Check your answers by turning to the Appendix, Section 1: Activity 3.

As mentioned earlier, the main difference between an electrochemical cell and a redox reaction is that in an electrochemical cell, the two half-reactions are separated so the electron transfer must occur through an external circuit. As scientists studied the cell, they also found it easier if they physically separated the half-reactions that were occurring so they could study each half-reaction independently.



To find out more about how and why the half-reactions are separated, read pages 394 and 395 of Section 13.2: Voltaic Cells in your textbook. Then complete the following questions.

4. a. Give two examples of a porous boundary.

   b. Why is it necessary that the boundary between the two electrolytes be porous?

   c. How is the porous boundary represented when writing the shorthand or cell notation?

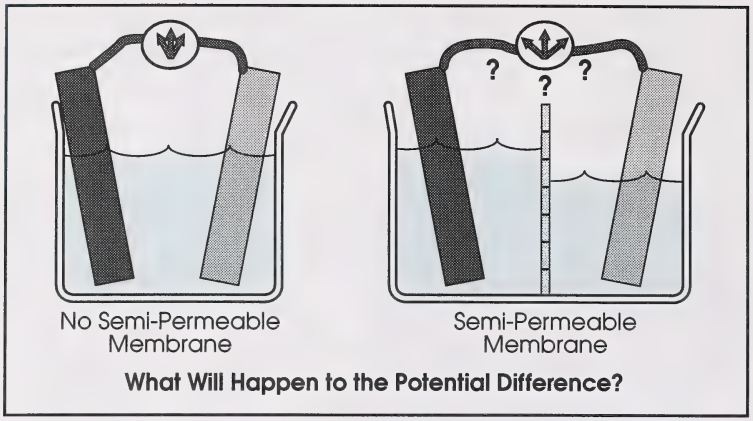
5. Why do chemists separate voltaic cells into two half-cells?

6. Match the following shorthand notations (a.-f.) to the descriptions of the electrochemical cells (i.-vi.).

Shorthand Notations	Cell Descriptions
a. $Sn_{(s)} \mid Sn^{4+}{}_{(aq)} \parallel Cu^{2+}{}_{(aq)} \mid Cu_{(s)}$	i. copper placed in a solution of copper(II) chloride and tin metal placed in a solution of tin(II) ions
b. $Mg_{(s)} \mid MgCl_{2(aq)} \parallel SnCl_{4(aq)} \mid Sn_{(s)}$	ii. a copper-magnesium cell
c. $Sn_{(s)} \mid SnCl_{2(aq)} \parallel CuCl_{2(aq)} \mid Cu_{(s)}$	iii. magnesium in a solution of magnesium chloride and tin in a solution of tin(II) chloride
d. $Mg_{(s)} \mid Mg^{2+}{}_{(aq)} \parallel Cu^{2+}{}_{(aq)} \mid Cu_{(s)}$	iv. a tin(IV) ion solution containing tin and a solution of magnesium ions containing magnesium
e. $Mg_{(s)} \mid Mg^{2+}{}_{(aq)} \parallel Sn^{2+}{}_{(aq)} \mid Sn_{(s)}$	v. two tin electrodes in solutions of tin(II) chloride and tin(IV) chloride respectively
f. $Sn_{(s)} \mid SnCl_{2(aq)} \parallel SnCl_{4(aq)} \mid Sn_{(s)}$	vi. copper placed in a copper(II) solution and tin placed in a tin(IV) solution

Check your answers by turning to the Appendix, Section 1: Activity 3.

You have now discovered what a porous cup does, but how does it affect the operation of a voltaic cell?



To study the effects of the porous barrier, perform the following investigation.

Investigation 13.3: Demonstration of a Voltaic Cell

Nelson CHEMISTRY

Read Investigation 13.3: Demonstration of a Voltaic Cell found on page 396 of your textbook. Pay special attention to the required components, safety aspects, and applied science skills.

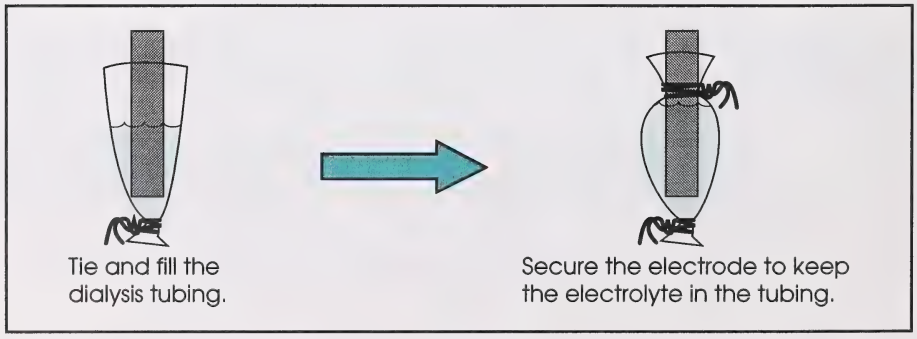
Science Skills

- ☐ A. Initiating
- ☑ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☐ F. Evaluating

Guidelines

The following changes can be made to the procedure:

- If a porous cup is not available, use the dialysis tubing supplied. Soak the dialysis tubing in water until you can open it up by rubbing it between your fingers; then, tie the bottom tightly with string or thread so it will not leak. Fill the dialysis tubing with one of your electrolytes. Insert your corresponding electrode, making sure you do not puncture the tubing, and tie the electrode into the bag. Store the electrolyte in a beaker to keep your lab area neat.



Important Note – Disposal Information

Any chemicals used in this module should not be poured down the sink, unless otherwise indicated. Keep waste materials in closed, labelled containers. Contact your nearest high school or fire department for disposal information.

- Use $Zn_{(s)}$, $Cu_{(s)}$, 0.10 mol/L $ZnSO_{4(aq)}$, 0.10 mol/L $CuSO_{4(aq)}$, and 0.10 mol/L $NaNO_{3(aq)}$ (salt bridge) as your electrodes and electrolytes (instead of $Ag_{(s)}$ and 0.10 mol/L $AgNO_{3(aq)}$).

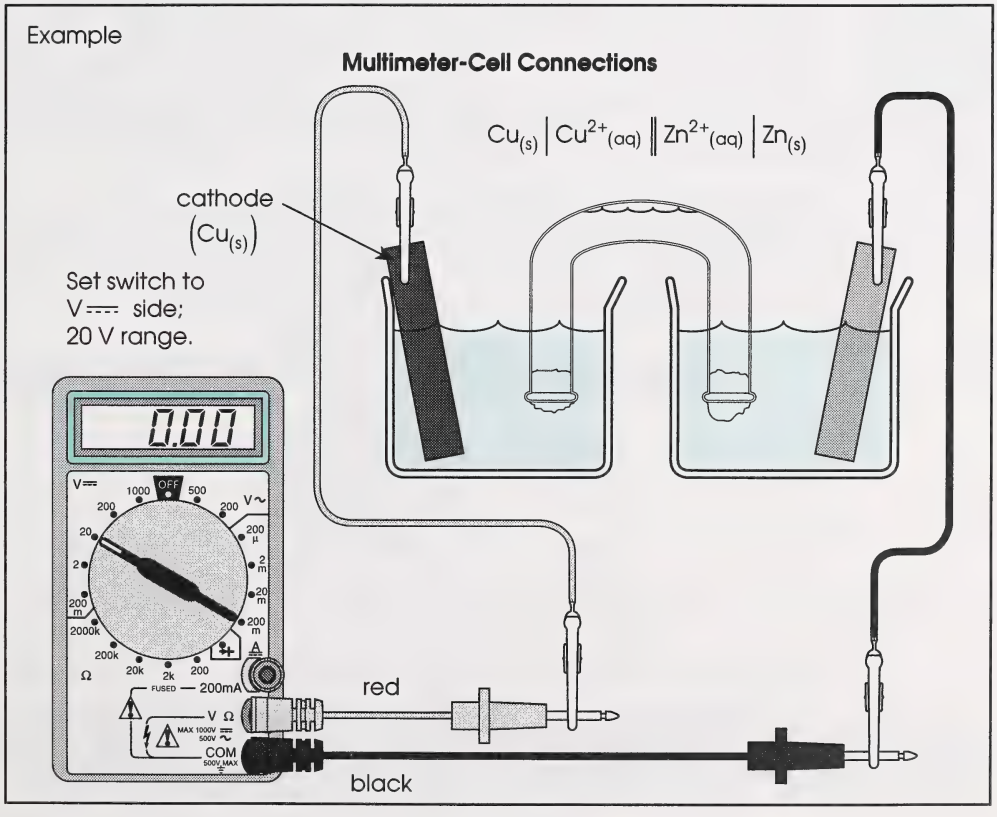
- Fill the U-tube with $NaNO_{3(aq)}$ and plug both ends with cotton batting.

Caution

Note: A paper towel soaked with $NaNO_{3(aq)}$ could be used instead of a U-tube, if necessary. Do not touch the paper towel with your hands. Use tweezers or tongs instead.

- Take your meter readings immediately upon hook-up.

Example

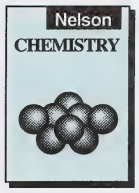


7. In addition to the qualitative observations you make, you should record your quantitative evidence in your notebook using a table with headings similar to the following:

potential difference – the difference in the amount of potential energy per unit charge between any two points

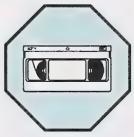
Type of Boundary	Cathode	Anode	Initial Potential Difference	Observations (after 5 min)	Observations (after 24 h)

8. Compare the values you recorded for **potential difference**.



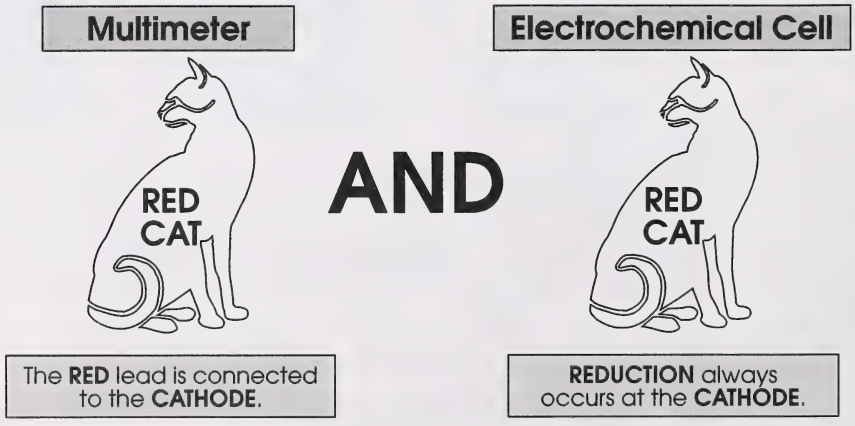
9. Why was the electrolyte in Figure 13.13 (a) on page 396 of your textbook made of $NaNO_{3(aq)}$ and not $Cu(NO_3)_{2(aq)}$ or $Zn(NO_3)_{2(aq)}$?

Check your answers by turning to the Appendix, Section 1: Activity 3.



You know that a porous membrane only helps keep the reactants separate, but how do the chemicals in a voltaic cell actually generate electricity? Watch video segment 2, Electrochemical Cells, from the series *Electrochemistry*; then, read the section A Theoretical Description of a Voltaic Cell beginning on page 396 of *Nelson Chemistry* and ending at the Exercise on page 399. The video and the reading will explain the basic principles of the electrochemical cell.

Notice that in the textbook example, copper is an anode, but in your previous investigation, it was a cathode. The most important thing to remember about what you saw and read is that **reduction occurs at the cathode**. If you remember this, then all of the other information you just studied must necessarily follow.



Answer the following questions to check your understanding of the basic principles of the electrochemical cell.

10. List the terms necessary to fill in the blanks.

    a. If reduction occurs at the cathode, then oxidation occurs at the _____.

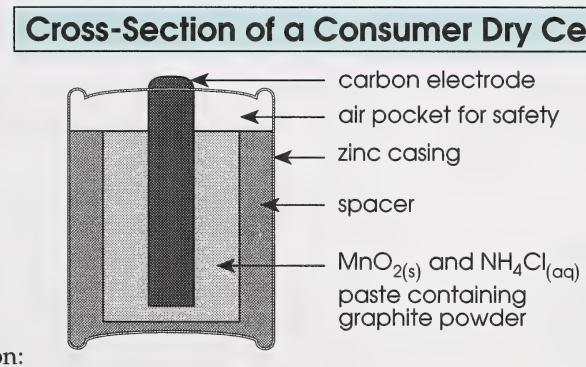
    b. Reduction means gain of electrons, so electrons must flow to the _____ from the _____. Electrons flow through the wire because that is how charge flows through metals. Charge can only flow through solution by using _____ to transfer the charge.

charge – a quantity of electricity; is measured in coulombs (C)

    c. Cations are usually reduced, so they must flow to the _____ to gain electrons. When cations are reduced, a metal usually forms which means the mass at the cathode usually _____. Likewise, the anode usually _____ mass because as oxidation occurs the metal electrode forms cations.

    d. When cations are reduced at the cathode, there become more anions than cations in the reduction half-cell. Since like charges repel, the anions will leave the _____ half-cell and migrate to the _____ half-cell. Likewise, the cations at the _____ must be in excess which means they will leave the _____ half-cell.

11. Using oxidation numbers and the balanced equation, identify the cathode, anode, porous membrane, species oxidized, and species reduced in the following diagram.



Cell Reaction:

$$Zn_{(s)} + 2\ MnO_{2(s)} + 2\ NH_4Cl_{(aq)} \rightarrow ZnCl_{2(aq)} + Mn_2O_{3(s)} + 2\ NH_{3(aq)} + H_2O_{(l)}$$

12. Answer textbook questions 16-21 on pages 399-400 of *Nelson Chemistry*.

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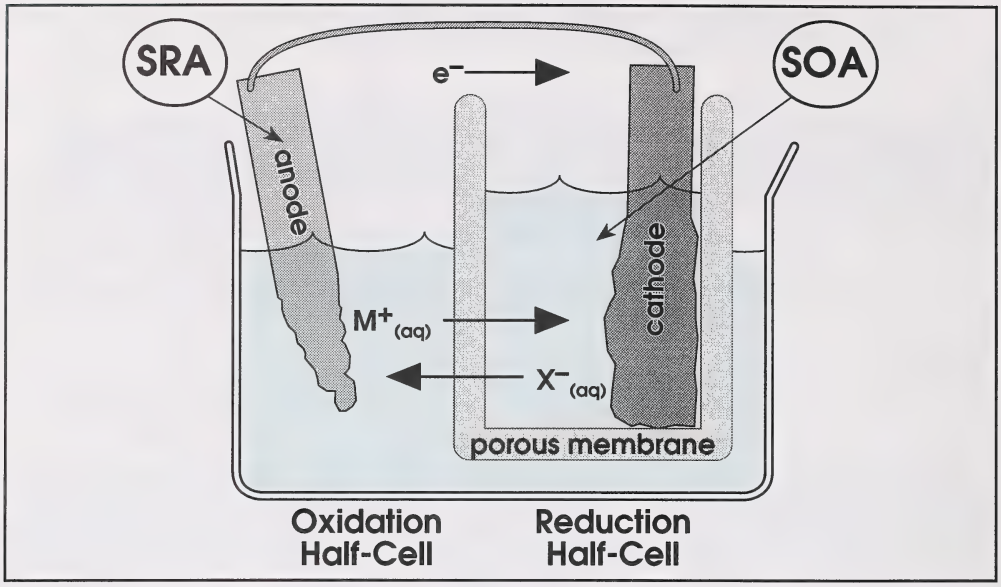
Check your answers by turning to the Appendix, Section 1: Activity 3.



Whenever chemical energy is converted into electrical energy, some type of voltaic cell has been constructed. The electrons released at the anode lose their energy as they move towards the cathode. Some of this energy is lost as heat, but the rest of the energy can be used to do work such as operating a cellular phone, a laptop computer, or your calculator.

In this activity, you learned how this energy is produced and how the parts of the voltaic cell efficiently work together. All voltaic cells operate the same way, regardless of their shape or the materials used in their construction. But before you can begin to explain how any of these cells operate, you must locate the cathode, because reduction always occurs at the cathode.

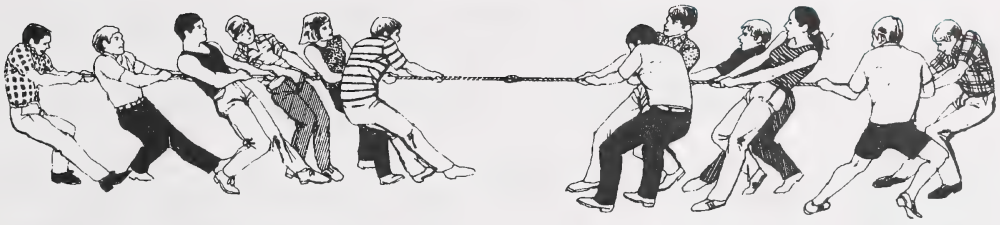
ACTIVITY 4

The Cell with Potential

Have you ever been involved in a tug-of-war?

1. With equal numbers on both sides of the tug-of-war rope, who usually wins?

Check your answers by turning to the Appendix, Section 1: Activity 4.



Electron transfer usually results from a tug-of-war. You have already seen that the competition for electrons is what distinguishes the strongest oxidizing and reducing agents. But how can you measure this competition? Can you put a number on the potential difference between an oxidizing agent and reducing agent?

In this activity, you will learn how scientists measure to what degree various substances like to tug on electrons in order to determine how easily oxidized and reduced they will be. Based on this, you will learn how to predict the potential difference of an electrochemical cell.

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First, however, you need to know how scientists went about setting the standards from which they could make the measurements that determined how much any particular substance wanted to be reduced. To find out, read pages 400-402 in your textbook stopping at the heading Measuring Standard Reduction Potentials. Then answer the following questions.

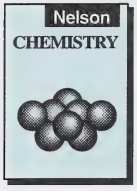
2. Match the following definitions to their correct chemical term. Chemical terms may be used more than once.

electrode potential – another name for reduction potential

Definitions	Chemical Terms
a. the half-reaction arbitrarily assigned an **electrode potential** of exactly zero volts	i. standard cell potential
b. a measure of how far apart the electrode potentials are of two standard half-cells ($\Delta E°$)	ii. reference half-cell
c. a measure of how far a reduction potential is from the standard reference half-cell	iii. standard reduction potential ($E°_r$)
d. the hydrogen half-cell – $2\ H^+_{(aq)} + 2\ e^- \rightarrow H_{2(g)}$ $\quad E°_r = 0.00$ V	iv. standard cell or half-cell
e. all substances at SATP conditions and, if in solution, are 1.0 mol/L	
f. the potential difference between two substances at standard conditions	

3. Why was it necessary to arbitrarily choose a standard reference half-cell?

4. Why does the choice of an arbitrary standard reference half-cell not affect a standard cell potential?



5. What significance do the positive and negative signs have on the $E°$ values on the Redox Half-Reactions table (page 552 of *Nelson Chemistry*)?

6. Answer textbook question 27 on page 405 of *Nelson Chemistry*.

Check your answers by turning to the Appendix, Section 1: Activity 4.

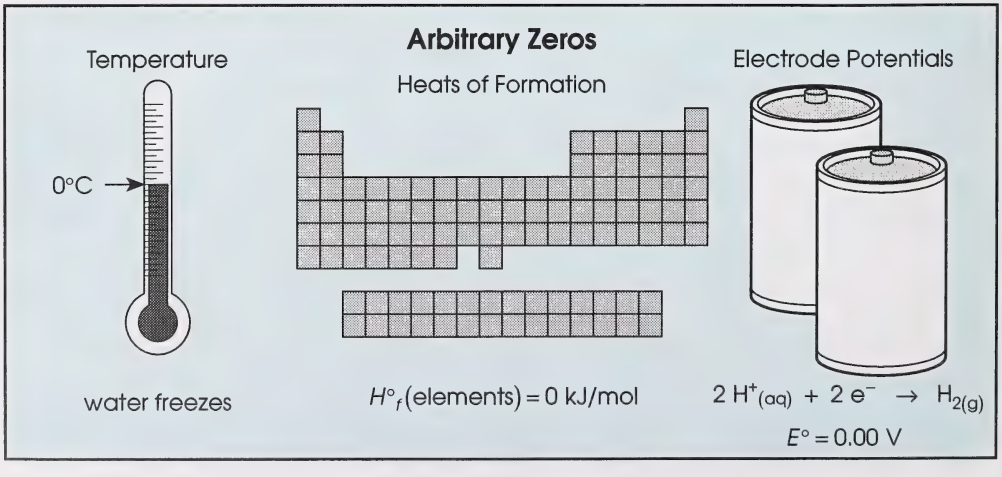
Important Note

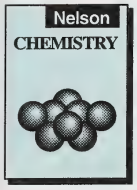
In your studies so far, you have seen **potential difference** represented by $E°$. You may also see **potential difference** represented by $V°$.

$$E° = V° \text{ and } \Delta E° = \Delta V°$$

The symbol $V°$ may represent **potential difference** on the Chemistry 30 Diploma Examination.

Since your textbook uses the symbol $E°$, however, this course will also continue using the symbol $E°$.





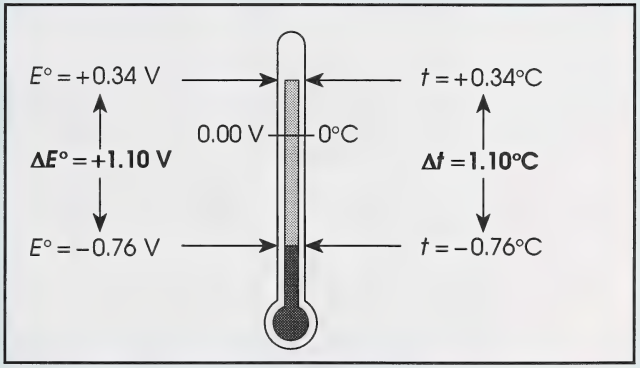
You now know how the standards were decided upon for electrode potentials, but how do scientists use these standards to predict what the potential difference of any particular electrochemical cell will be? To find out, read the information in the section Measuring Standard Reduction Potentials found on pages 402-404 of your textbook.

From what you read, you may have noticed that when predicting the cell potential, you are just finding out how far apart the two half-reactions are, just like finding out how far apart two temperature readings on a thermometer are. If the strongest oxidizing agent (SOA) is higher than the strongest reducing agent (SRA), the cell potential (potential difference) is positive. The formula for potential *difference*, $\Delta E^\circ = E^\circ_r(\text{cathode}) - E^\circ_r(\text{anode})$, is based on this idea.

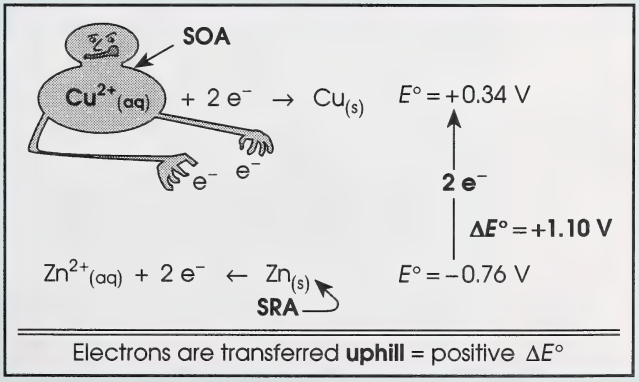
Example

In a $Zn_{(s)} \mid Zn^{2+}_{(aq)} \parallel Cu^{2+}_{(aq)} \mid Cu_{(s)}$ cell, you could use your table of redox half-reactions as follows:

Step 1 To determine the cell potential, see how far apart the reduction potentials are, just as you would if you were finding a temperature difference.

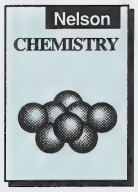


Step 2 A spontaneous reaction (SOA is higher than SRA on redox table) tells you ΔE° is positive.



Therefore, you know the cell potential for a standard copper-zinc cell is +1.10 V.

To help you understand the concepts of cell potential, strongest oxidizing agent, and strongest reducing agent even further, answer the following.



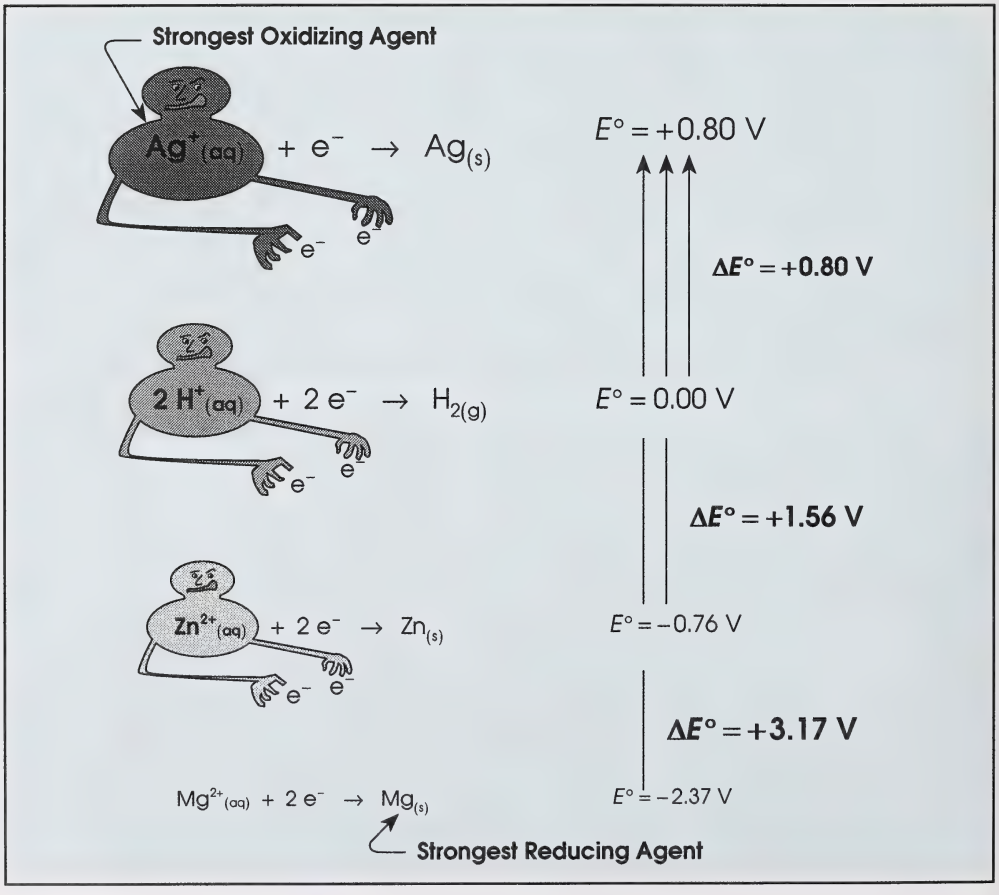
7. Answer textbook questions 23, 24, 25, and 26 on page 405 of *Nelson Chemistry*.

8. When you have to multiply a half-reaction by a number so the electrons balance, you do NOT multiply the $E°$ value. Why?

9. Answer textbook questions 17, 19, and 20 on page 422 of *Nelson Chemistry*.

10. (Enrichment – Optional) Complete the prediction of Lab Exercise 13C: Series Cells on page 406 of *Nelson Chemistry*.

Check your answers by turning to the Appendix, Section 1: Activity 4.


Strongest Oxidizing Agent
Ag+(aq) + e− → Ag(s)
E° = +0.80 V
ΔE° = +0.80 V
2 H+(aq) + 2 e− → H2(g)
E° = 0.00 V
ΔE° = +1.56 V
Zn2+(aq) + 2 e− → Zn(s)
E° = −0.76 V
ΔE° = +3.17 V
Mg2+(aq) + 2 e− → Mg(s)
E° = −2.37 V
Strongest Reducing Agent

An alternate method to using the formula $\Delta E° = E°_r(\text{cathode}) - E°_r(\text{anode})$ for finding the standard cell potential follows next. In this alternate method, the $E°$ values are written with the half-reaction equations. The sign of the $E°$ value for the oxidation half-reaction is reversed.

Example

For the following standard cell, $Zn_{(s)} \mid Zn^{2+}_{(aq)} \parallel Cu^{2+}_{(aq)} \mid Cu_{(s)}$, write the half-reactions, the overall net ionic equation, and calculate the cell potential.

Answer

anode	OA	SOA	cathode
$Zn_{(s)}$	$Zn^{2+}_{(aq)}$	$Cu^{2+}_{(aq)}$	$Cu_{(s)}$
SRA			RA

$Cu^{2+}_{(aq)} + 2\,e^- \rightarrow Cu_{(s)}$ $\qquad E°_r = +0.34$

$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2\,e^-$ $\qquad E°_{ox} = \boxed{+0.76}$

add to get

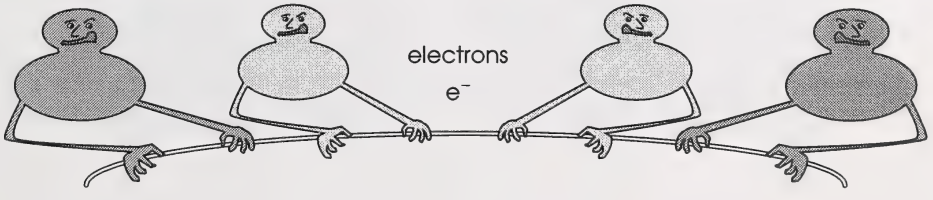
$Cu^{2+}_{(aq)} + Zn_{(s)} \rightarrow Cu_{(s)} + Zn^{2+}_{(aq)}$ $\qquad \Delta E° = +1.10$

Notes

The difference in this example is that the sign of the $E°$ value for the oxidation half-reaction (Zn) is **reversed**, just as the half-reaction itself is.

The $E°$ values are then **added**, just as the half-reactions are, to get the net cell potential. Everything else is done the same as the preceding formula method.

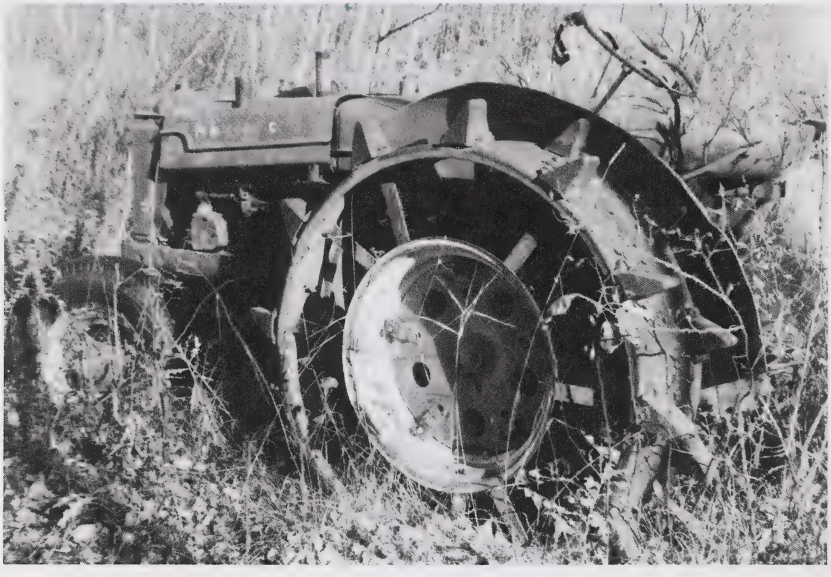
In this activity, you discovered that the competition for electrons is much like a tug-of-war. The stronger the oxidizing agent, the higher its reduction potential, so the more easily it can steal electrons from a reducing agent. The difference between the potential energies of the electrons will determine the net cell potential (potential difference). It is from these relative potential differences of substances that the table of redox half-reaction equations was made.



ACTIVITY 5

Application and Evaluation

How do you stop rust?



Rust accounts for billions of dollars of damage every year, so it's no wonder that people try to develop new ways to prevent it from happening. Some of these ways are based on the concepts of electrochemistry.

corrosion – the oxidation of any material, especially metals

Answer the next two questions to explore different ways of preventing corrosion, and to use your knowledge of chemistry to make a practical decision about corrosion.

1. Which of the following will NOT protect iron from corrosion?

   a. paint
   b. plastic coating
   c. make an alloy
   d. rust remover
   e. oxide coating
   f. metal plating
   g. attach a more reactive metal
   h. add salt
   i. run a current through it

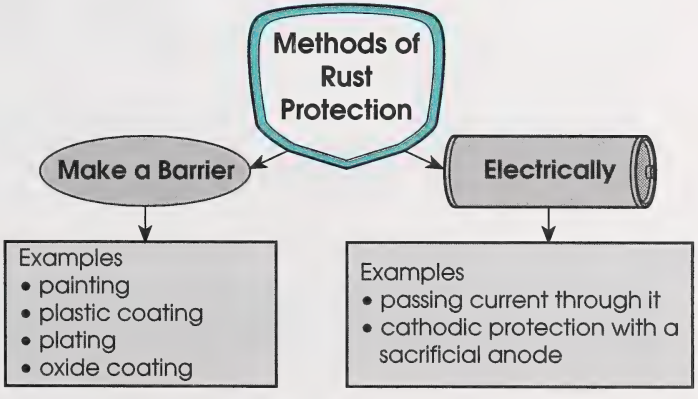
2. You are trying to put an iron mesh onto your screen door and are trying to decide what type of fasteners to use. One friend says to use brass because brass doesn't rust. Another says to use zinc. Yet another says aluminum, and still another friend says that the regular iron ones will work fine. Based on what you have learned so far about redox reactions and voltaic cells, predict which fastener(s) would be best to use. Which fastener(s) should not be used?

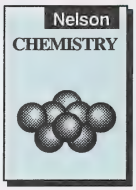
Check your answers by turning to the Appendix, Section 1: Activity 5.

In the preceding question you used the same principles as those used to protect boat motors and pipelines from rusting, as well as to protect such things as ships' hulls from corroding. The method involves making a voltaic cell using a **sacrificial anode** attached to the metal you want to protect. In order for it to be oxidized first, the sacrificial anode must be a stronger reducing agent than the metal you want to protect. The metal you want to protect then becomes the cathode. Since reduction occurs at the cathode, corrosion (oxidation) cannot occur. The process of protecting metal in this way is called **cathodic protection**.

sacrificial anode – a metal that is oxidized to protect another metal from corrosion. It is a stronger reducing agent than the protected metal.

cathodic protection – the method of forcing a metal to become a cathode by attaching a stronger reducing agent to it





Since rusting and corrosion are really oxidation processes, scientists often use oxidation potentials when discussing cathodic protection. To learn how oxidation potentials are used, read the section Oxidation Potentials found on pages 404-405 of your textbook and answer the next three questions.

3. Use your table of redox half-reactions to help you decide whether the following statements are true or false. Correct the false statements to make them true.

   a. $PbSO_{4(s)} + 2\,e^- \rightarrow Pb_{(s)} + SO_4^{2-}{}_{(aq)}$ $\quad E° = +0.36$ V

   b. $Cd_{(s)} \rightarrow Cd^{2+}{}_{(aq)} + 2\,e^-$ $\quad E° = +0.40$ V

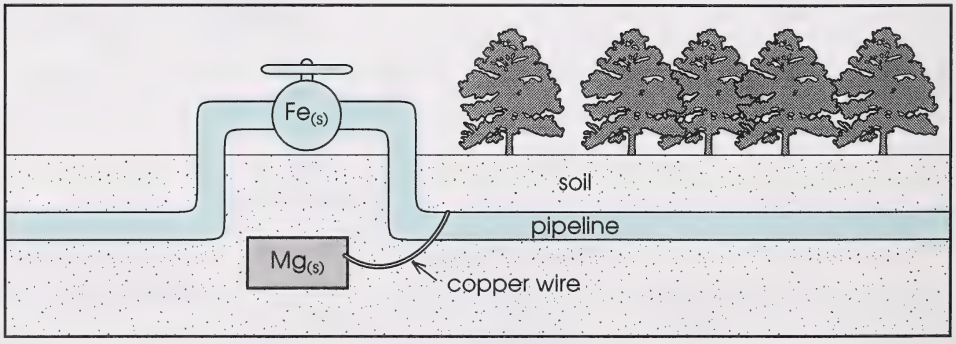
   c. The oxidation potential for aluminum producing aluminum ions is +1.66 V.

   d. When bromine becomes bromide, the potential of the half-reaction is –1.07 V.

4. Based on the following clues, rank and list the hypothetical metals M, P, Q, and T from highest oxidation potential to lowest oxidation potential.

   - M is more easily oxidized than Q.
   - T is a sacrificial anode for P.
   - P will cathodically protect Q.
   - When M and T touch, T is the cathode.

5. In gas pipelines, blocks of magnesium are often connected to valves by conductive wire to prevent the valves from corroding. Using the following diagram as a guide, identify the cathode, anode, and electrolyte. Also label the direction of flow of the electrons.



Check your answers by turning to the Appendix, Section 1: Activity 5.

Throughout this section, you have been studying the basic principles of spontaneous electrochemical cells, how they relate to batteries, and how these same principles can be used in other applications such as metal protection. You should now be able to build and label the parts of a voltaic cell given the appropriate materials. In addition, you should be able to predict the potential difference of the cell you construct.

Investigation 13.4: Testing Voltaic Cells

Nelson CHEMISTRY

Do Investigation 13.4: Testing Voltaic Cells on page 407 of your textbook to give yourself an opportunity to confirm what you have learned about electrochemical cells. Pay special attention to the required components, safety aspects, and applied science skills.

Science Skills

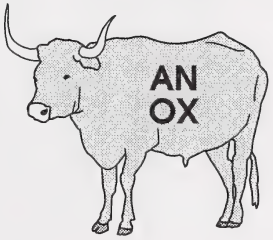
- [x] A. Initiating
- [x] B. Collecting
- [x] C. Organizing
- [x] D. Analysing
- [ ] E. Synthesizing
- [x] F. Evaluating

Guidelines

- Keep the following two hints in mind:

REDUCTION always occurs at the **CAT**HODE.

AND



At the **AN**ODE, **OX**IDATION occurs.

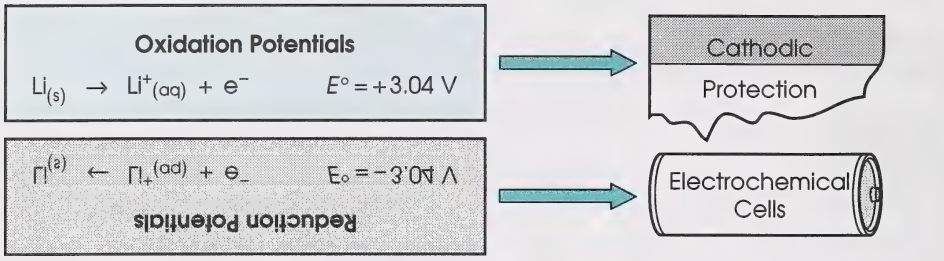
- Use a U-tube containing $NaNO_{3(aq)}$ or dialysis tubing for your porous membrane.
- You will not be using the $Ag_{(s)}$ and 0.10 mol/L $AgNO_{3(aq)}$.
- Build at least three cells.
- Use your multimeter, set to direct current volts (V ⎓), 20 V range, as your voltmeter.
- Use alligator clips and wires to connect the multimeter wires to the cell electrodes, as you did in Investigation 13.3.

6. Make a prediction of the cell potential for each cell you will construct.
7. Write down the steps in your procedure.

8. Construct an evidence table, perform the procedure, and record your evidence.

9. If you did not get the same results as you predicted, explain why they are different.

Check your answers by turning to the Appendix, Section 1: Activity 5.



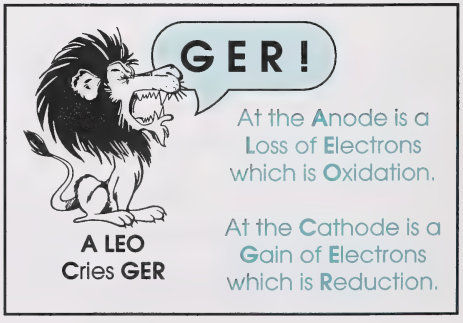
In this activity, you learned that besides separating redox half-reactions and making electrochemical cells, redox reactions can be used to protect metals from corrosion by cathodic protection. In cathodic protection, corrosion such as rust can be stopped by a simple electrochemical cell formed between two different metals. Scientists can use half-cells to predict the cell potential of such cells. Oxidation potentials are sometimes used to choose appropriate sacrificial anodes for use in this method.

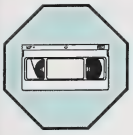
Follow-up Activities

If you had difficulties understanding the concepts in the activities, it is recommended that you do the Extra Help. If you had a clear understanding of the concepts, it is recommended that you do the Enrichment.

Extra Help

The main thing to remember when working with voltaic cells is that the half-reactions of a simple redox reaction are separated so that reduction occurs at the cathode and oxidation occurs at the anode. Electron transfer then must occur through an external circuit.

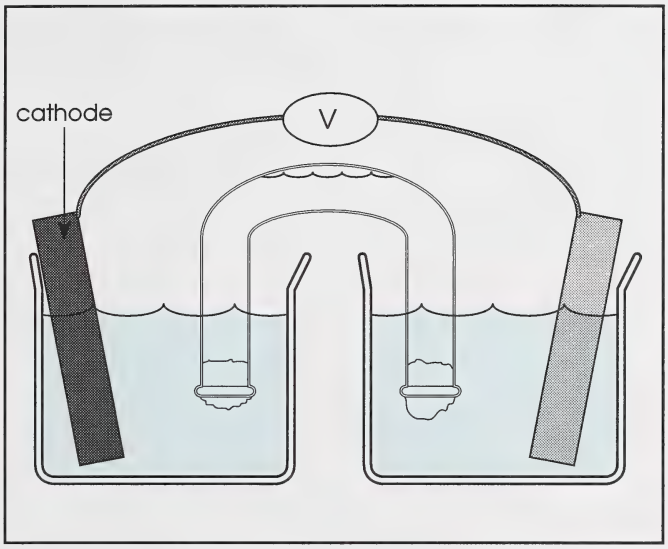

GER!
At the Anode is a Loss of Electrons which is Oxidation.
At the Cathode is a Gain of Electrons which is Reduction.
A LEO
Cries GER



To better understand the relationship between the cell and potential difference, watch video segment 3, Designing Electrochemical Cells, from the series *Electrochemistry*. Complete questions 1, 2, and 3 as you watch this video.

1. Reproduce the following diagram. Use the given terms to label your diagram.

- anode
- cation flow
- gains mass
- reduction
- anion flow
- loses mass
- oxidation
- electron flow
- salt bridge



2. Copy the following paragraph and fill in the blanks with the following words or phrases (some may be used more than once):

- cathode
- anode
- strongest oxidizing agent
- electrode potentials
- positive
- electrolyte
- strongest reducing agent
- net cell potential
- circuit
- spontaneous
- energy

Voltaic cells consist of three main parts – the _____, _____, and _____. In the cell, electrons move from the _____ to the _____. The solution between the two electrodes is an _____, because it conducts electricity and is needed to complete the _____. As the electrons move through the wire, they lose _____. The amount of energy they lose depends on how far apart the _____ are. The further apart they are, the greater the _____ of the cell is. If the _____ is higher than the _____ on a table of reduction half-reactions, then the reaction is _____ and the net cell potential is _____. An example of this type of cell can be used to start a robot.

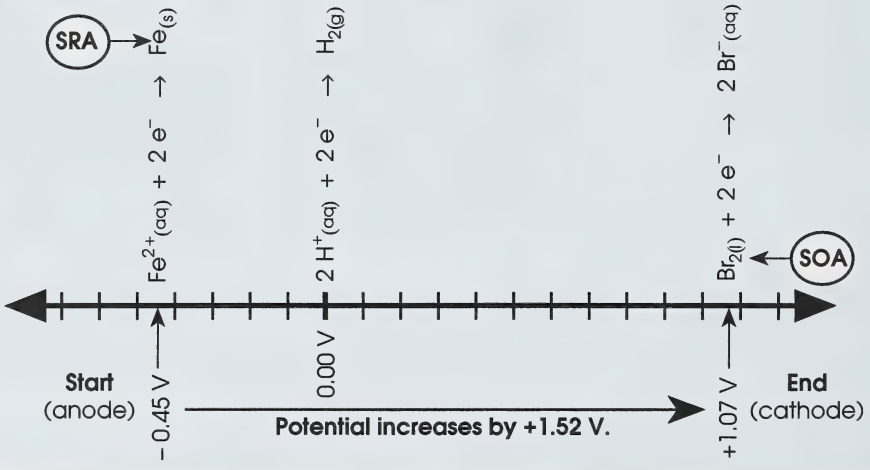
Hint: By turning your table of reduction half-reactions sideways like a number line, you can determine the cell potential quickly by calculating how many units you move, starting at the anode.

Example

Determine the cell potential of the following cell:

$$Fe_{(s)} \left| Fe^{2+}_{(aq)} \right\| Br_{2(l)}, Br^{-}_{(aq)} \left| C_{(s)} \right.$$

First of all, the strongest oxidizing agent is $Br_{2(l)}$ and the strongest reducing agent is $Fe_{(s)}$. Therefore, carbon is the cathode, and iron is the anode. From the number line, you can see that $\Delta E°_{net} = +1.52$ V.



3. Match the cells on the left with their correct net cell potentials on the right.

a. $Al_{(s)} \left| Al^{3+}_{(aq)} \right\| Pb^{2+}_{(aq)} \left| Pb_{(s)} \right.$

b. $Cr_{(s)} \left| Cr^{2+}_{(aq)} \right\| Fe^{2+}_{(aq)} \left| Fe_{(s)} \right.$

c. $Pt_{(s)} \left| O_{2(g)}, OH^{-}_{(aq)} \right\| F^{-}_{(aq)}, F_{2(g)} \left| Pt_{(s)} \right.$

d. $Ni_{(s)} \left| Ni^{2+}_{(aq)} \right\| H^{+}_{(aq)}, H_{2(g)} \left| C_{(s)} \right.$

i. $\Delta E°_{net} = +0.26$ V

ii. $\Delta E°_{net} = +1.53$ V

iii. $\Delta E°_{net} = +2.47$ V

iv. $\Delta E°_{net} = +0.46$ V

As you can see from the previous question, every redox reaction involves an energy transfer, many in the form of simple voltaic cells. The further the two half-reactions are apart, the more energy per unit charge the battery can deliver. Besides batteries, another example of a simple voltaic cell in action is in cathodic protection.

4. To keep her iron eavestrough from corroding in the atmosphere (air + water), Sylvia attaches little pieces of magnesium to the eavestrough. Explain whether or not this will cathodically protect the roof.

Check your answers by turning to the Appendix, Section 1: Extra Help.

Enrichment

Complete **one** of the following.

1. First there was the flashlight. Next the transistor radio. Then the cellular phone and the laptop computer. As society makes more and more demands for portability, the demand for portable power increases. But as technology improves, alternative sources of power such as solar cells are also being developed.

   Do your own research or talk to a battery producer to find various ways in which chemical batteries are important to society, and predict what the future may hold for them. One possible source of information is The Answer Centre, TransAlta Utilities, Calgary, Alberta.

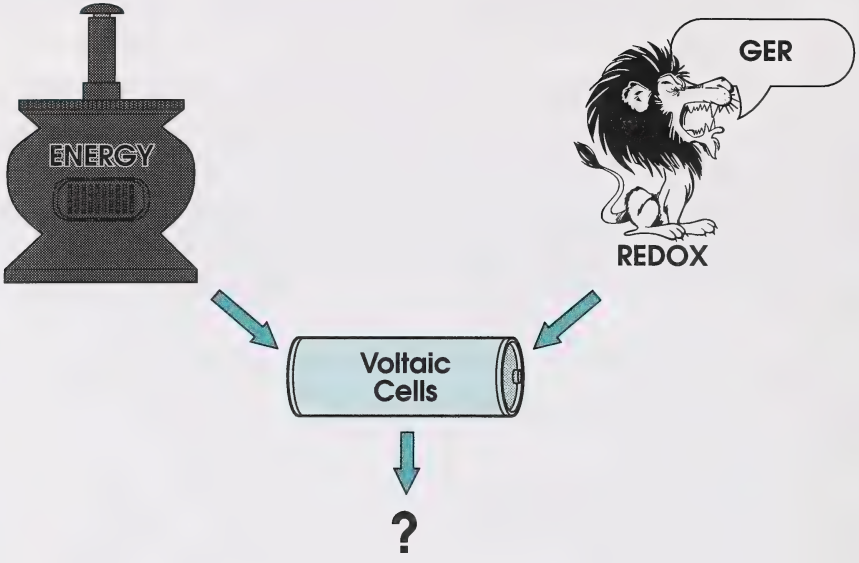
2. In 1992, it is estimated that more than 1.5 billion Ni-Cad rechargeable batteries and even more disposable batteries were sold. As a result, the question of recycling batteries occurs again and again.

   Visit the Industrial Waste Branch of the Alberta government, or a place such as a garage that collects used batteries, and find out what happens to the batteries once they are dropped off. Focus on the methods of battery disposal that may be used in your area. In addition, you may want to suggest some alternatives that may be used and the implications these involve.

3. What is the real difference between carbon batteries, alkaline batteries, and rechargeables in terms of their performance? To find out, devise an experiment that would test the life of these three types of batteries.

Check your answers by turning to the Appendix, Section 1: Enrichment.

Conclusion



If you would like additional review of the concepts discussed in this section and have access to the video *Electrochemical Cells* (Volume 1, *CHEM Study* series – Ward's Natural Science Ltd.), watch it now. As in this section, the video discusses the basic components of a system called the electrochemical cell, and how the potential difference in a cell is determined.

You have seen the concepts you learned in Module 3 utilized and expanded upon in electrochemical cells. This section has also introduced the enormous potential for application of these concepts as well as the possible detriments that accompany them.

40

Section 1 Assignment: Chemical Energy at Work

Review the Evaluation information found in the introductory pages of this module.

It is important to number and clearly identify each page with the following information at the top:

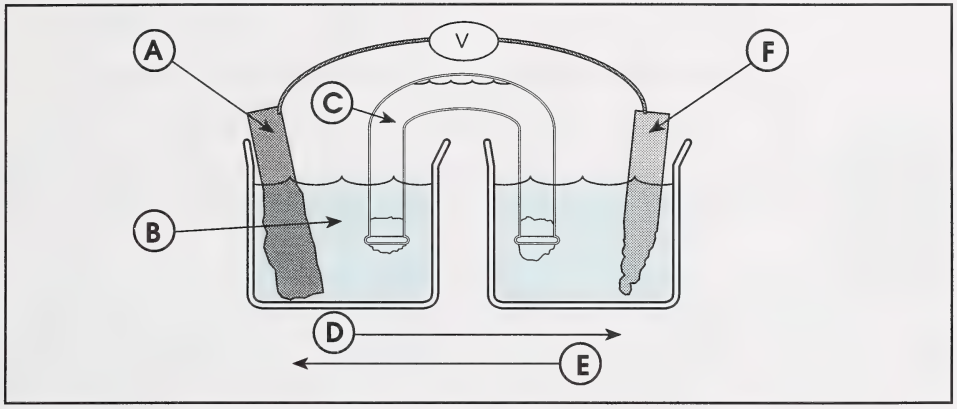
Chemistry 30 – Module 4 Section 1 Assignment Page # Name and ID#

Be sure to write legibly. Leave a wide left margin and number all of your pages.

1. The following diagram represents a zinc/silver electrochemical cell with the following notation:

$$Zn_{(s)} \left| Zn(NO_3)_{2(aq)} \right\| AgNO_{3(aq)} \left| Ag_{(s)} \right.$$

The salt bridge contains $NaNO_{3(aq)}$. Use the preceding notation and the diagram to answer the seven parts of this question.



Response Page Setup

a. Copy the preceding diagram onto your assignment page and label the appropriate letters with the following terms (not all letters may be used and the same letter may be used more than once):

- cathode
- location of strongest oxidizing agent (SOA)
- location of strongest reducing agent (SRA)
- direction of anion flow
- salt bridge

(5 marks)

b. The ions initially present in the solution indicated by point (B) would be _____. **(1 mark)**

c. The ions initially present in the salt bridge would be _____. **(1 mark)**

d. The ions that would migrate through the salt bridge from each solution would be _____. **(2 marks)**

e. The direction of flow of electrons would be towards the electrode labelled letter _____. **(1 mark)**

f. Describe what happens to the concentrations of the following ions as the cell operates: $Ag^{+}_{(aq)}$, $Zn^{2+}_{(aq)}$, $Na^{+}_{(aq)}$, and $NO_3^{-}{}_{(aq)}$. **(4 marks)**

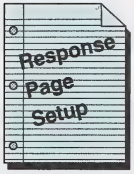
g. Calculate the standard cell potential of this silver/zinc cell. **(2 marks)**

2. Calculate the potential difference of the cells represented by the following information:

   a. $Cd_{(s)} \mid Cd^{2+}{}_{(aq)} \parallel Ag^{+}{}_{(aq)} \mid Ag_{(s)}$ **(2 marks)**

   b. $Cu_{(s)} \mid Cu^{2+}{}_{(aq)}, NO_3^{-}{}_{(aq)} \parallel Na^{+}{}_{(aq)}, SO_4^{2-}{}_{(aq)} \mid Pb_{(s)}$ **(2 marks)**

3. If $F_{2(g)} + 2\,e^- \rightarrow 2\,F^-{}_{(aq)}$ was chosen as the standard reference electrode, explain how it would affect the answer to question 2. a. **(2 marks)**



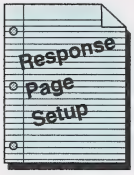
4. Rearrange the following sets of half-reactions to create an electrochemical cell in each case. Write the labelled cathode half-reaction, anode half-reaction, and net ionic reaction. Also calculate the net cell potential. Show all your work clearly. Marks will be given for both your work and your answer. **(6 marks)**

   a. $Fe^{3+}{}_{(aq)} + 3\,e^- \rightarrow Fe_{(s)}$ $\quad E° = -0.04$ V

   $F_{2(g)} + 2\,H^+{}_{(aq)} + 2\,e^- \rightarrow HF_{(aq)}$ $\quad E° = +3.03$ V

   b. $NO_3^{-}{}_{(aq)} + 2\,H^+{}_{(aq)} + e^- \rightarrow NO_{2(g)} + H_2O_{(l)}$ $\quad E° = +0.80$ V

   $SO_4^{2-}{}_{(aq)} + 4\,H^+{}_{(aq)} + 2\,e^- \rightarrow H_2SO_{3(aq)} + H_2O_{(l)}$ $\quad E° = +0.17$ V



5. Use the following information to answer all parts of this question. Show all your work clearly. Marks will be given for both your work and your answer.

   > A voltaic cell was constructed from two half-cells as follows:
   >
   > One half-cell was made from an unknown metal and a solution containing its ion.
   >
   > The other half-cell was $Cu_{(s)} \mid Cu(NO_3)_{2(aq)}$. The net cell potential was +1.13 V.
   >
   > After 24 hours, it was observed that the blue colour of the copper half-cell solution had intensified.

   a. Determine the reduction potential of the unknown metal. **(2 marks)**

   b. Considering experimental error, use your table of redox half-reactions to predict the most likely identity of the unknown metal. **(1 mark)**

Science Skills

- ☑ A. Initiating
- ☐ B. Collecting
- ☐ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☐ F. Evaluating

6. A fishing enthusiast has just purchased a fishing trawler (a large boat) and wants to cathodically protect the iron bottom with another metal. She has two unlabelled grey metals from which to choose – one is zinc; one is silver. Knowing that only one of these grey metals will suit her purpose, she sets out to identify each grey metal by performing an experiment. Besides a sample of iron and the grey metal samples, the only other materials available are salt water, a paper towel, copper wire, a multimeter with wire leads, and two pails.

   a. Which of the two metals, zinc or silver, would suit her purpose? Explain. **(2 marks)**

   b. Write a brief procedure that describes how the two grey metals could be identified, including predicted results. **(3 marks)**

Response Page Setup

Reproduce the science skills box after your response to this question. Remember to indicate your evaluation of your skill level for each identified science skill. Leave a space for the teacher assessment.

Self:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐
Teacher:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐

7. Three standard half-cells are constructed from unknown metals $X_{(s)}$, $Y_{(s)}$, and $Z_{(s)}$ and their ions $X^{2+}_{(aq)}$, $Y^{2+}_{(aq)}$, and $Z^{2+}_{(aq)}$. Different combinations of the three half-cells are tried when constructing various electrochemical cells. The following information is obtained from the different combinations:

   - An $X_{(s)} \mid X^{2+}_{(aq)}$ standard half-cell attracts anions when the other standard half-cell is $Y_{(s)} \mid Y^{2+}_{(aq)}$. The $\Delta E° = 0.50$ V.
   - In a standard cell between $X_{(s)}$ and $Z_{(s)}$, $X_{(s)}$ is the cathode and $\Delta E° = 0.14$ V.

Response Page Setup

Determine the cathode, anode, and net cell potential of a standard cell between $Y_{(s)} \mid Y^{2+}_{(aq)}$ and $Z_{(s)} \mid Z^{2+}_{(aq)}$. Show all your work clearly. Marks will be given for both your work and your answer. **(4 marks)**

SECTION 2

Turning Electricity into Chemistry

Coins. You use them to pay for CDs, to make purchases from vending machines, to make phone calls, and maybe even to make a wish. You might even collect them. But how are they made? Are they all made in the same way? And how are they related to the batteries you discussed in the previous section?

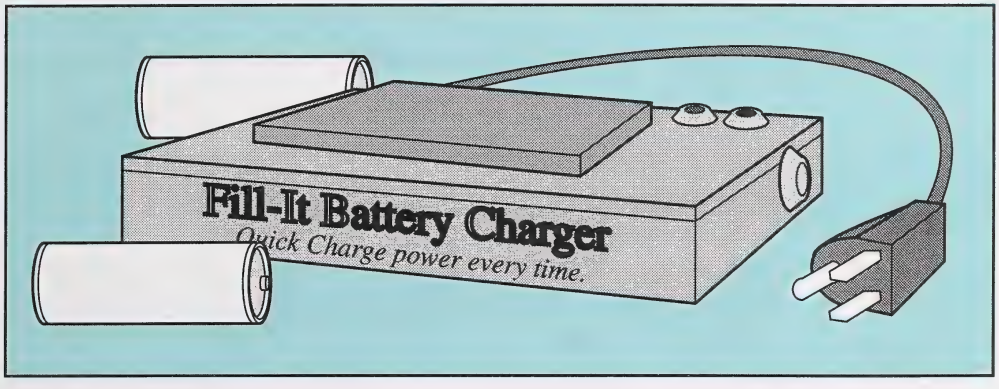
In Module 3, you were introduced to the basic concepts of electrochemistry through a brief look at metallurgical processes. How do the metallurgical processes involved in coin making relate to the concept of electroplating you will examine in this section?

You will discover the answers to these types of questions as you study the process of converting electrical energy into chemical energy. You will construct electric cells to demonstrate these processes. Some common examples will be studied to give you an idea of how prevalent electrochemistry is in your life. Finally, you will perform chemical calculations related to such processes as electroplating.

ACTIVITY 1

Electrolytic Cells

You may often hear the statement, "If you use it, put it back when you're finished."



How does the preceding statement apply to rechargeable batteries? What are you doing when you recharge batteries? Before answering this question, you should review the basic principles behind a voltaic cell. Use Figure 4.2 as you answer the next three questions to test your understanding.

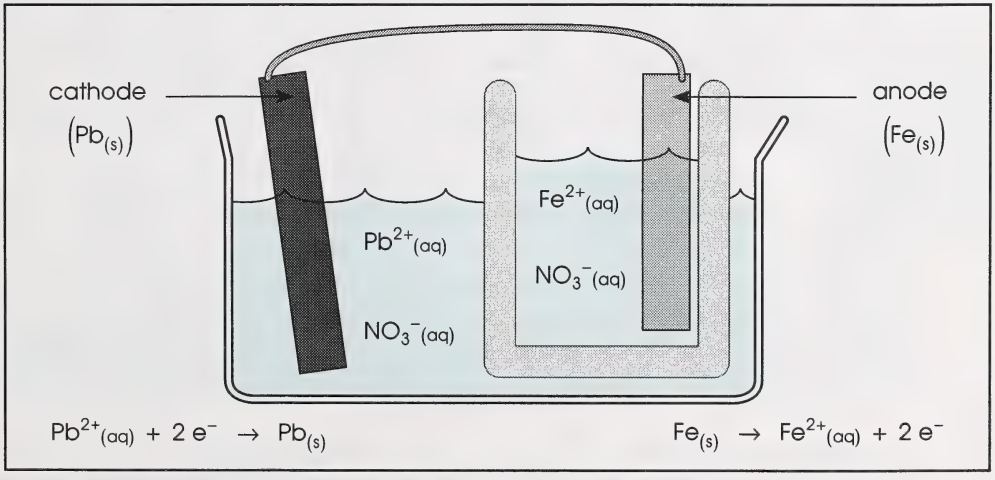


FIGURE 4.2: A Voltaic Cell

1. Determine whether the cell would still operate if each of the following changes were made. Assume you start with the voltaic cell from Figure 4.2 each time. Explain your answer for each.

   a. The $Fe(NO_3)_{2(aq)}$ is replaced with $NaNO_{3(aq)}$.

   b. The iron anode is replaced with a copper electrode.

   c. The porous cup is removed and the two solutions are allowed to mix.

2. If the iron anode in Figure 4.2 is replaced with a copper electrode, the porous cup would no longer be required. Why?

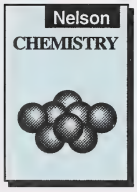
3. Answer the following multiple-choice question.

electrolytic cell
a non-spontaneous cell that converts electrical energy into chemical energy

If the cell in question 2 was forced to run, this would then be an example of an electrolytic cell. One major difference between a voltaic cell and an electrolytic cell (as described in question 2) would most likely be that an electrolytic cell must have

A. no porous membrane
B. a positive potential difference
C. a non-spontaneous electrochemical reaction
D. no anode

Check your answers by turning to the Appendix, Section 2: Activity 1.



To find out more about what an electrolytic cell is and how it operates, read the section Electrolysis on pages 407-408 of your textbook, ending after Table 13.4 on page 408. Then answer the next three questions.

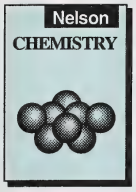
4. Prepare a chart with the given headings in your notebook. List as many similarities and differences as possible between a voltaic cell and an electrolytic cell.

COMPARING THE VOLTAIC CELL WITH THE ELECTROLYTIC CELL

Similarities	Differences	
	Voltaic	Electrolytic

5. a. Besides having a non-spontaneous reaction occur in an electrolytic cell, the main difference between a voltaic cell and an electrolytic cell is that the electrolytic cell requires an _____.

   b. One way to recognize whether a cell is voltaic or electrolytic is by looking up the reactants in your table of redox half-reactions. If the strongest oxidizing agent is at the same level or lower than the strongest reducing agent, then the cell is an _____ cell. Another way is to see if the cell has an external _____.



6. Answer textbook questions 13 (on page 421) and 21 (on page 422) in *Nelson Chemistry*.

Check your answers by turning to the Appendix, Section 2: Activity 1.

power – the rate at which energy is being consumed in watts (W); 1 watt = 1 joule of energy being used in 1 s

Did you notice that it is the non-spontaneous reactions that require an external power source? That is because they require energy to force the reaction to occur.

But how does the need for external energy affect the way in which electrolytic cells are labelled? Study Figure 4.3, paying special attention to the differences between the cells.

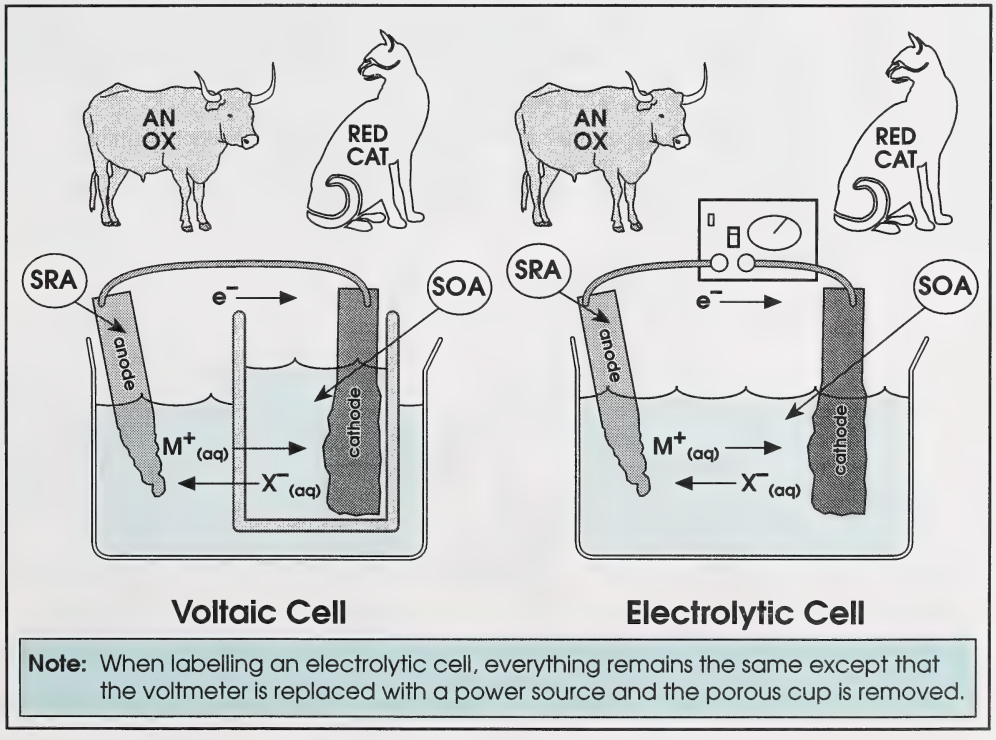
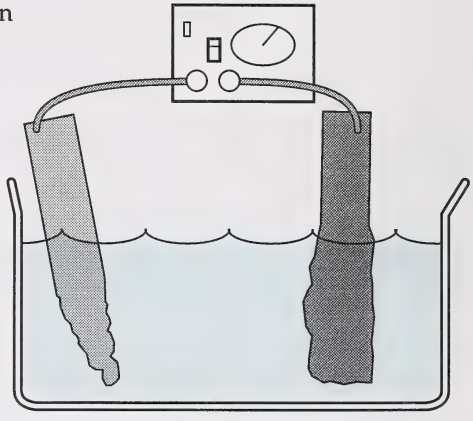


FIGURE 4.3: Labelling Voltaic and Electrolytic Cells

As you can see from Figure 4.3, the cathode is the electrode to which the electrons move. The power source acts as a pump to force the electrons away from the anode and towards the cathode. This forces reduction to happen at the cathode and oxidation to happen at the anode.

7. The following two half-reactions occur in an electrolytic cell:



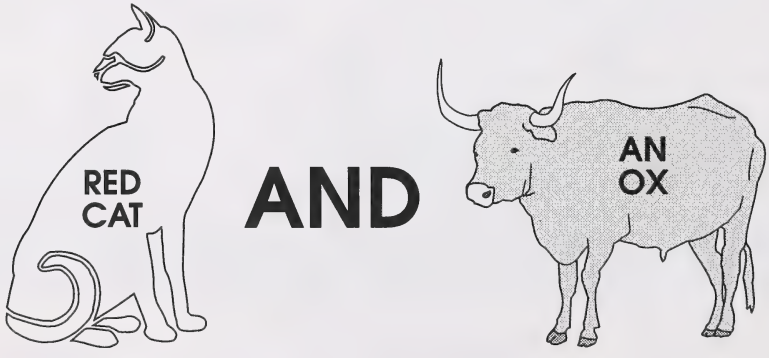
$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2\,e^-$

$Sn^{2+}_{(aq)} + 2\,e^- \rightarrow Sn_{(s)}$

If the cathode is a carbon rod and the anode is copper, label a diagram similar to the one on the right, indicating the anode, cathode, power supply, electrolyte, $Cu_{(s)}$, $C_{(s)}$, $Sn(NO_3)_{2(aq)}$, electron flow, anion flow, cation flow, strongest reducing agent, and strongest oxidizing agent.

Check your answers by turning to the Appendix, Section 2: Activity 1.

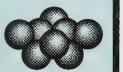
As you can see from the previous questions, there are not many differences between a voltaic cell and an electrolytic cell. Most importantly, however, is to remember the basic principles that **all** cells are based on.



REDUCTION always occurs at the **CAT**HODE.

At the **AN**ODE, **OX**IDATION occurs.

Nelson CHEMISTRY



Science Skills

- ☐ A. Initiating
- ☑ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☐ F. Evaluating

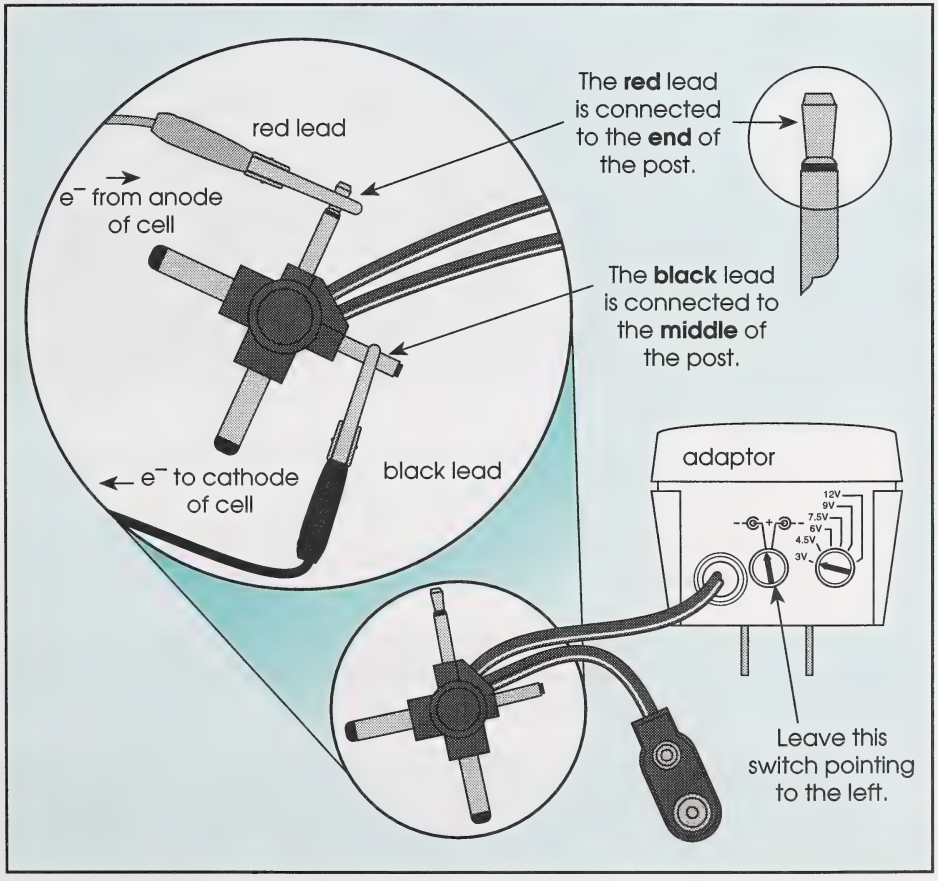
Caution

Investigation 13.5: A Potassium Iodide Electrolytic Cell

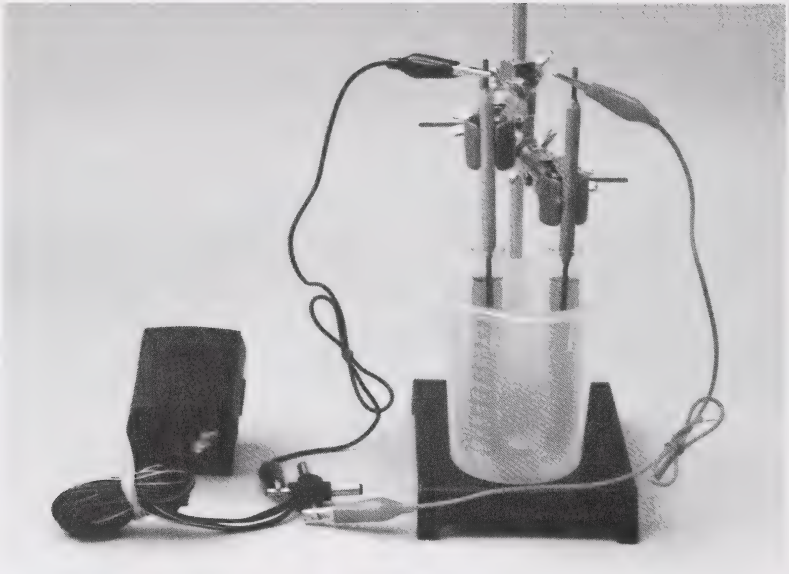
Perform Investigation 13.5: A Potassium Iodide Electrolytic Cell found on pages 408-409 of *Nelson Chemistry*, in order to see how the basic principles of a non-spontaneous electrolytic cell work. Pay special attention to the required components, safety aspects, and applied science skills.

Guidelines

- CAUTION! Do not plug your power supply into the wall until your experiment has been set up and all of your initial observations have been recorded.
- For your power supply, you can use the low voltage universal AC adapter and connect your leads as shown in the following diagram. Make sure the voltage is set on 3 V.

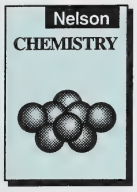


- Pencil leads can be used for electrodes in your electrolytic cell (as in Investigation 13.2 described in Section 1, Activity 2).
- Use 20 mL of 0.10 mol/L $KI_{(aq)}$ in a U-tube (supported by a 250 mL beaker) instead of a petri dish.
- Do not do the trichlorotrifluoroethane test if you do not have access to this organic solvent and supervised laboratory facilities.
- Your equipment should be set up similar to the following photograph.



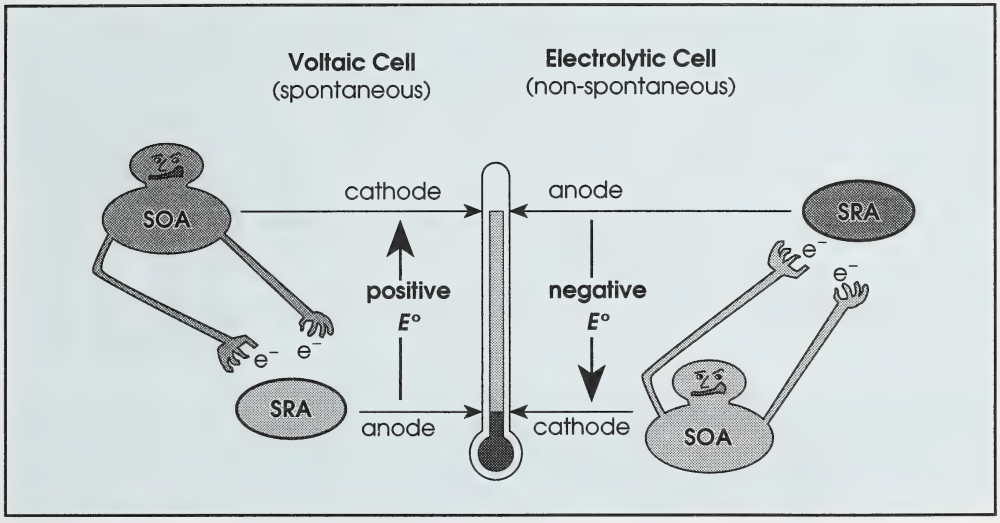
8. Prepare an evidence table, perform the investigation according to the given procedure, and record your observations.
9. Based on the evidence you obtained, what products were formed at the anode and cathode?
10. Using your table of redox half-reactions, write the oxidation and reduction half-reaction equations and the net ionic equation that occurred during the electrolysis of $KI_{(aq)}$. (The species present are $K^+_{(aq)}$, $I^-_{(aq)}$, and $H_2O_{(l)}$.)

Check your answers by turning to the Appendix, Section 2: Activity 1.

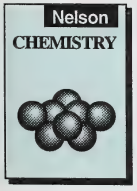


To get a further explanation of what went on in Investigation 13.5, and to learn how to estimate the voltage required to produce the iodine, read the section The Potassium Iodide Electrolytic Cell: A Synthesis on pages 409-410 of *Nelson Chemistry*.

As you may have noticed, the method for calculating net cell potentials is exactly the same as before. You simply start at the anode half-reaction and determine how far the cathode half-reaction is away from it. A negative potential difference means a non-spontaneous reaction. The equation $\Delta E° = E°_r(\text{cathode}) - E°_r(\text{anode})$ also still holds, just as it did for the voltaic cell. This is shown in the following diagram.



Practise calculating the approximate minimum cell potential required for a reaction, by completing the next two questions.



11. Complete textbook questions 28 and 29 on page 410 of *Nelson Chemistry*.

12. Using a Ni-Cad battery as an example, explain how a secondary cell acts as both a voltaic cell and an electrolytic cell. Useful information can be found in Table 13.2 on page 392 of your textbook.

Check your answers by turning to the Appendix, Section 2: Activity 1.

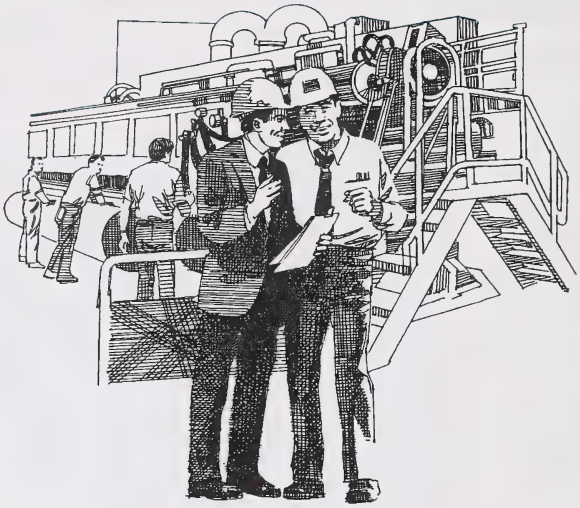
In this activity, you discovered the similarities and differences between a voltaic cell and an electrolytic cell. Both types of cell meet the criteria for a basic electric cell – two electrodes and an electrolyte. The main difference is that the voltaic cell is a system that changes chemical energy into electrical energy while an electrolytic cell is a system that changes electrical energy into chemical energy.

ACTIVITY

2 Electrolysis – Synthesis and Prediction

Imagine a life without a chemical as simple as sodium hydroxide. What would it be like? Normally you don't think about how **electrolysis** is important in your life, but electrolysis is used in many ways to help produce the chemicals you need to make some of the things you often take for granted. Every time you use a loonie, for instance, you are making use of a product that was produced by an electrolytic process.

electrolysis – the process of forcing a non-spontaneous redox reaction to occur using electrical energy



DID YOU KNOW?

Paper, rayon, soap, bleach, dyes, drugs, and even some foods all rely on caustic soda as one of the ingredients necessary for their production. Caustic soda is another name for *sodium hydroxide*, and sodium hydroxide is largely produced by electrolysis.

Nelson CHEMISTRY

In this Activity, you will look more closely at how electrolysis can be used to produce various compounds. Investigation 13.6: Demonstration of Electrolysis on page 411 of *Nelson Chemistry* will give you some insight into the basic process involved in the production of various chemicals by electrolysis.

Investigation 13.6: Demonstration of Electrolysis

Science Skills

- ☑ A. Initiating
- ☑ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☑ F. Evaluating

Read Investigation 13.6: Demonstration of Electrolysis. Pay special attention to the required components, safety aspects, and applied science skills.

Pre-lab

1. a. Predict the anode, cathode, and net reactions, including $E°$ values, for each of the electrolytic cells indicated in the problem.

   b. Looking at each half-reaction, include what evidence should be seen at each electrode to confirm your predictions.

PATHWAYS

If you have access to supervised laboratory facilities, do Part A.
If you do not have access to laboratory facilities, do Part B.

Part A

This investigation may be done as a demonstration by your teacher, or you may choose to do the investigation yourself.

2. Construct an appropriate evidence table, perform the investigation, and record your evidence.

Check your answers by turning to the Appendix, Section 2: Activity 2.

As your analysis and evaluation, do the questions following Part B.

End of Part A

Part B

Guidelines

Change the experimental design as follows:

- Carry out each electrolysis in a U-tube containing 20 mL of solution, as you did in Investigation 13.5.
- To make each solution, dissolve about 2 mL (1/2 teaspoon) of solid in about 25 mL water.

  If the copper(II) sulphate solution turns somewhat cloudy when the solid is added to water, it should not appreciably affect your results.

- Use two pencil leads for your electrodes as you did in Investigation 13.5. Rinse the electrodes after each use.

- Use your universal adapter as your power supply. Do NOT plug it in until your experiment has been set up and all of your initial observations have been recorded.
- Adjust the setting to 3 V on your universal adapter. If you do not get any noticeable results within two minutes, change the setting to 4.5 V. The increased voltage should produce results.
- Use litmus paper as an aid to identify each half-reaction that occurred in the electrolysis.

Materials

- 2 mL each of
  - copper(II) sulphate pentahydrate($CuSO_4 \bullet 5H_2O_{(s)}$)
  - sodium sulphate ($Na_2SO_{4(s)}$)
  - sodium chloride ($NaCl_{(s)}$)
- water
- U-tube
- retort stand with 2 clamps
- 250 mL beaker
- 100 mL beaker
- universal adapter
- 2 wire leads
- 2 electrodes (pencil leads)

3. Write the steps of the procedure you will follow.

4. Construct an appropriate evidence table, perform the investigation, and record your evidence.

Check your answers by turning to the Appendix, Section 2: Activity 2.

Do the following questions as your analysis and evaluation.

End of Part B

5. As your analysis, compare your experimental results with your predictions.

6. As your evaluation, suggest at least two changes that could be made to improve your investigation.

Check your answers by turning to the Appendix, Section 2: Activity 2.

anomaly – an unexpected occurrence

In the performance of the preceding investigation, you probably encountered what scientists encounter in research and industry quite often – an **anomaly**. Your $NaCl_{(aq)}$ electrolytic cell should have given you the same results as the $Na_2SO_{4(aq)}$ cell. However, you encountered evidence of a substance smelling like bleach. This substance turned litmus paper white. What do you think it was?

If you used only your sense of smell, you might say the substance was chlorine; and you would be right. But how did chlorine get produced in the cell at the anode? To find out, follow the logic of the next few questions.

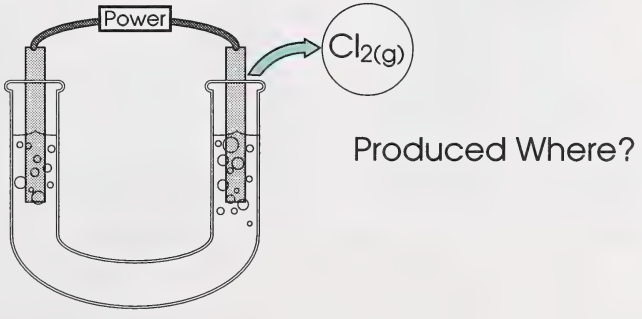
7. Since the chlorine $\left(Cl_{2(g)}\right)$ was produced at the anode, what process – oxidation or reduction – produced the chlorine?

8. What starting material was present in the solution from which the chlorine could have been produced?

9. Propose a possible half-reaction equation to illustrate the production of the $Cl_{2(g)}$.

10. Find the chlorine half-reaction equation on your redox table. Considering how close the chlorine half-reaction is to the water half-reaction on the table, propose a possible explanation as to how the chlorine could have been produced in your cell.

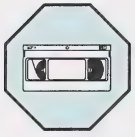
Check your answers by turning to the Appendix, Section 2: Activity 2.



The anomaly you discovered in the preceding investigation has its practical uses. Chlorine gas is produced by a similar electrolytic process by companies such as Dow Chemical Canada Inc. in Fort Saskatchewan, Alberta. The chlorine is used to make bleach, hydrochloric acid, and other useful substances.

Important Note

There is something important you should realize. Anomalies or exceptions such as the chlorine half-reaction you discovered **DO NOT** change the way you predict redox reactions. You will still continue to predict redox reactions the same way you always have.



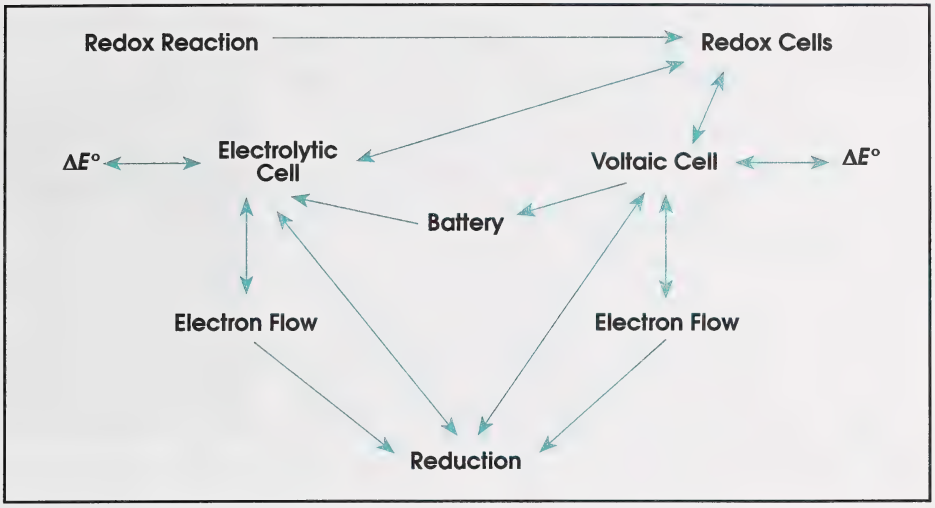
So far, you have seen how electrolysis is done. But how, exactly, does it work from a chemical standpoint? Watch video segment 6, Electroplating, of the series *Electrochemistry*, to see how electrolysis occurs chemically. Then answer the next three questions to test your understanding of what you have watched.

11. Copy and complete the following table in your notebook:

Method of Metal Protection	Description	Advantages	Disadvantages
plating with less reactive metal			
plating with more reactive metal			
galvanizing			
cathodic protection			

12. On each arrow in the following web, write a word or a short statement from the list that shows how the two connected terms are related. Some statements may be used more than once.

- spontaneous reaction
- negative
- separate half-reactions
- a possible power source
- gains (collects) electrons
- non-spontaneous reaction
- positive
- more than one
- anode to cathode
- occurs at cathode



13. a. Reproduce Figure 4.4 in your notebook. Identify and label the following parts on your diagram:

- electrolytic cell
- power source
- the cathodes
- cation flow
- voltaic cell
- flow of electrons
- the anodes
- anion flow

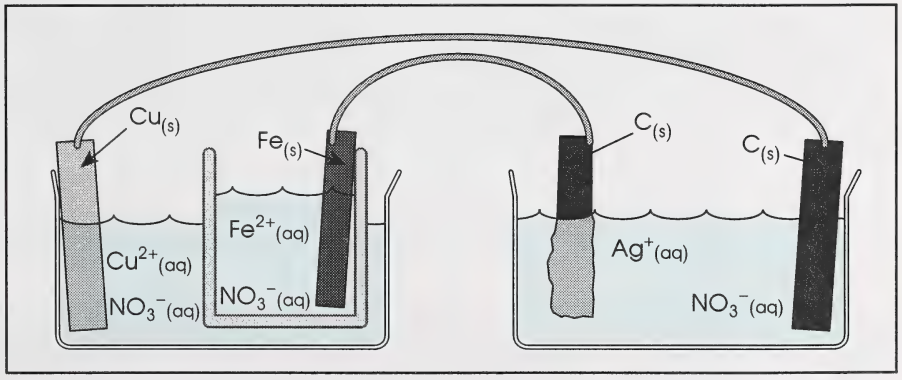


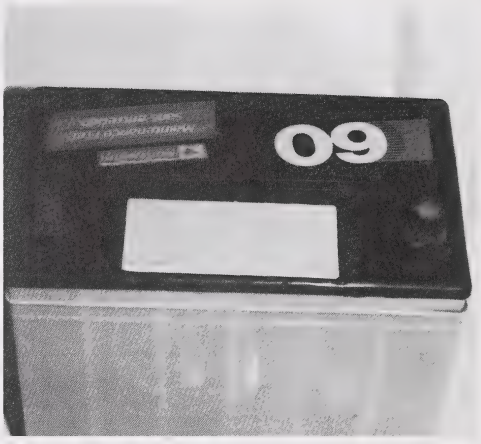
FIGURE 4.4: Cells

b. Based on the given information, will a current flow through the wires in the setup in Figure 4.4? Explain.

Check your answers by turning to the Appendix, Section 2: Activity 2.

You have been introduced to several applications of electrolysis techniques. One application that you may be familiar with is in the operation of your car and its battery. Have you ever wondered why your car battery doesn't *go dead* as long as your car is running, but as soon as you leave your headlights on for a while when the car is turned off, the battery dies?

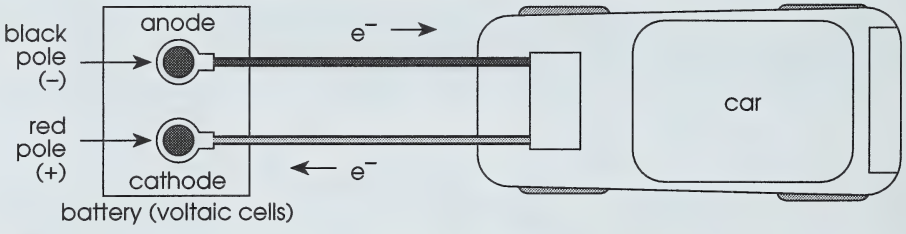
At the beginning of this module, you were introduced to the lead-acid battery, the type used in most automobiles, as a collection of secondary or rechargeable cells. What does rechargeable mean?



DAVE MORRIS

Your car battery converts chemical energy into electrical energy to **start** your car. The battery is operating as a series of voltaic cells at this time.

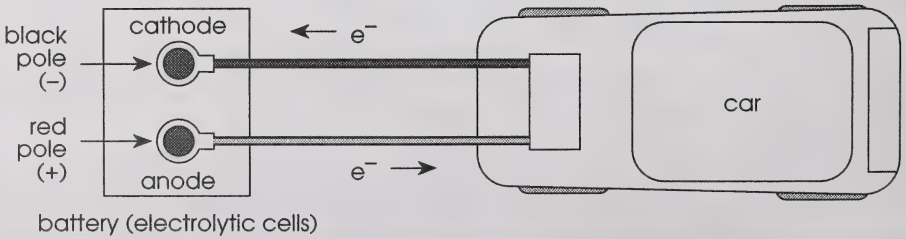
Starting the Car (Battery Is Discharging)

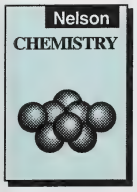


Note: The anode and cathode are defined based on the chemical process occurring.

However, once your car is running, the mechanical energy of the motor is transformed into electrical energy and runs your car battery as a series of electrolytic cells. This reverses the reactions in your battery, converting electrical energy into chemical energy, thus recharging the battery.

Running Car (Battery Is Recharging)





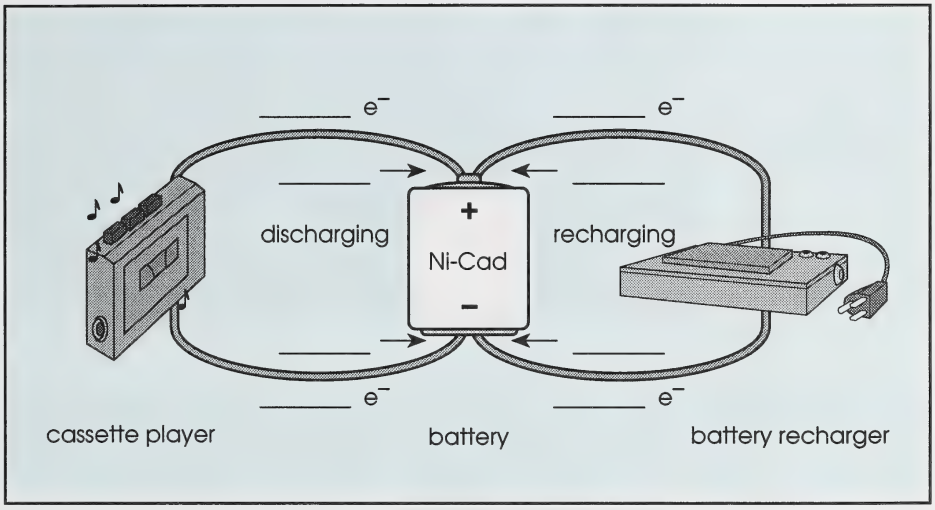
All rechargeable batteries work on a similar principle. The half-reactions involved in a lead-acid cell (car battery) are described in Table 13.2 on page 392 of your textbook. Use this information, your redox table, and the preceding discussion to answer the following questions.

14. Write the anode, cathode, and net cell reaction for a lead-acid cell when it is discharging (voltaic cell). Include cell potentials.

15. Write the anode, cathode, and net cell reaction for a lead-acid cell when it is recharging (electrolytic cell). Include cell potentials.

16. Compare the two sets of reactions – the discharging and recharging cells – and their net cell potentials.

17. The following diagram represents a nickel-cadmium (Ni-Cad) rechargeable cell when it is discharging and recharging. Label the anode, cathode, and electron flow in each case.



Check your answers by turning to the Appendix, Section 2: Activity 2.

In this activity, you practised predicting the products of electrolysis, and saw how electrolytic cells and voltaic cells can interact. While these systems are quite similar, they also have some fundamental differences. The main difference is that the voltaic cell is a system that can be used to change chemical energy into electrical energy and the electrolytic cell is a system that will not operate on its own because it needs a power source to change electrical energy into chemical energy.

ACTIVITY

3 STS and Electrolysis

1. Look at the following photograph. What is one thing that the aluminum foil has in common with the apple juice? What does the foil have in common with the jewellery?

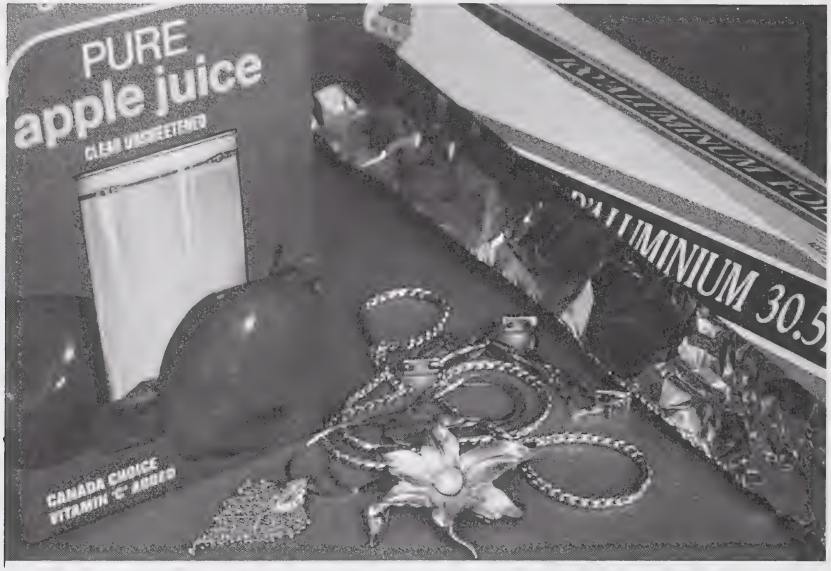
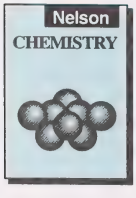


PHOTO SEARCH LTD.

Check your answers by turning to the Appendix, Section 2: Activity 3.

electroplating – covering a material with a thin layer of protective metal by the process of electrolysis

What did you identify as a common element? One thing in common between the aluminum and the jewellery that is relevant to this activity is that both were most likely produced by some form of electrolysis. Electrolysis and **electroplating** of metals is widespread in industry and is very important economically to many aspects of society.



In this activity, you are going to study some of the ways in which electrolysis is used in society. One of the uses of electrolysis is in the production of elements. For a brief description of this use, read pages 411-414 (omit page 412) of *Nelson Chemistry*, starting at the heading Science and Technology of Electrolysis and ending at the Exercise. Then answer the next four questions to check your understanding.

2. Fill in the blanks with the appropriate words.

   Cations always move to the _____. Since cations are _____ ions, when the cation of a binary ionic compound is reduced at the cathode, the element it forms is a _____. Anions, on the other hand, form _____ elements at the anode if they undergo _____.

3. Which of the following $E°$ values do not match the half-reaction it is written for under standard conditions? Explain your choices. (**Hint:** What do phases have to do with the half-reaction?)

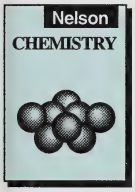
   a. $K^{+}_{(aq)} + e^{-} \rightarrow K_{(s)}$ $\quad E° = -2.93$ V

   b. $Ni^{2+}_{(l)} + 2\,e^{-} \rightarrow Ni_{(l)}$ $\quad E° = -0.26$ V

   c. $2\,Cl^{-}_{(l)} \rightarrow Cl_{2(g)} + 2\,e^{-}$ $\quad E° = -1.36$ V

   d. $Au_{(l)} \rightarrow Au^{3+}_{(l)} + 3\,e^{-}$ $\quad E° = -1.50$ V

4. Answer textbook questions 30, 31, and 32 from pages 414-415 of *Nelson Chemistry*.



5. Write equations for the cathode, anode, and net cell reactions that occur when the following substances are electrolysed.

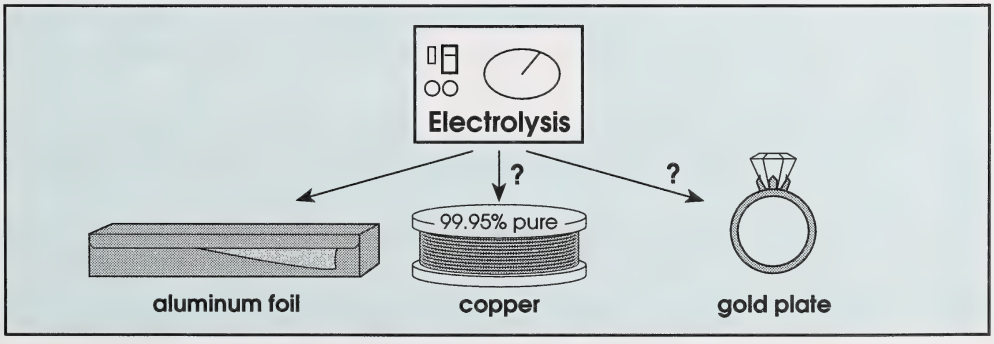
   a. $Ni_2O_{3(l)}$

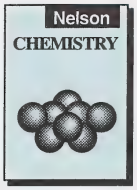
   b. $FeO_{(l)}$

   c. $NaCl_{(l)}$

   d. $Pb(NO_3)_{2(aq)}$

Check your answers by turning to the Appendix, Section 2: Activity 3.





You now know how the aluminum in aluminum foil is obtained, but how is the gold on a gold-plated piece of jewellery obtained? or the extremely pure copper needed for copper wire? The sections Refining of Metals and Electroplating found on pages 415-417 of your textbook will explain two other ways electrolysis is applied to meet society's needs. Read these sections and answer the next three questions to learn more about these processes.

WESTFILE INC.

6. Refer to Figure 13.26 on page 415 of *Nelson Chemistry* to answer this question.

   In the purification of metals, match each of the items in the left column with its location in an electrolytic cell.

Item	Location
a. impure metal	i. bottom of cell under anode
b. pure metal	ii. bottom of cell under cathode
c. molten electrolyte	iii. not used in process
d. solution of metal ion	iv. anode
e. sludge	v. cathode
f. no precipitate	vi. cell

7. Decide whether each of the following statements about electroplating is true or false.

   a. For some metals, the effectiveness of electroplating depends on which anions are in solution.

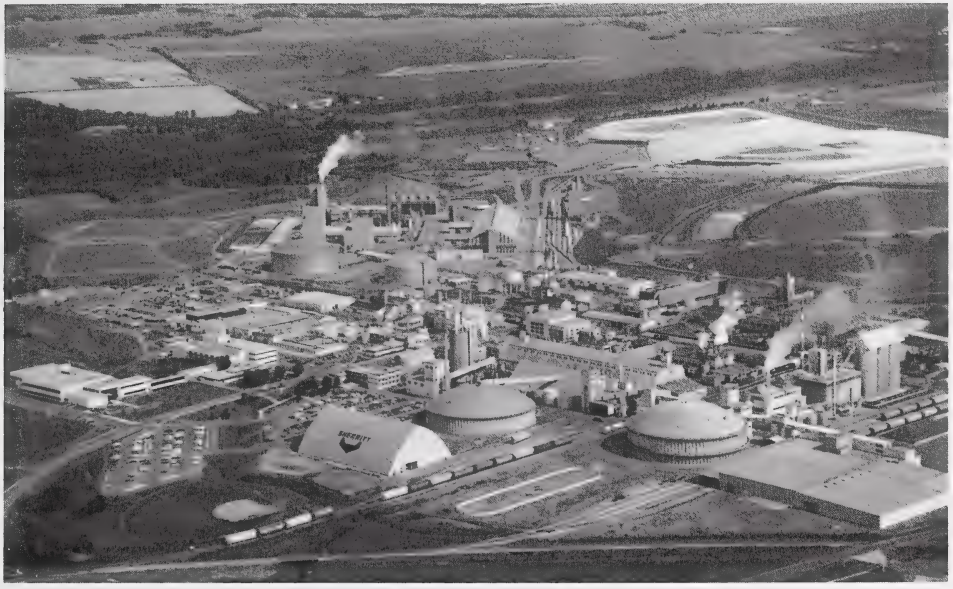
   b. One possible reason for poor adhesion of one metal on another is that the metal attached to the cathode had an oxide coating that was not removed first.

   c. Technological problem solving plays a major role in improving electroplating processes. Scientific knowledge mainly assists in the process.

   d. The metal to be electroplated should be placed at the anode, because that is where the cations are forced to go.

Nelson CHEMISTRY

8. Read page 412 of your textbook to discover the evolution of the loonie, as produced now by Sherritt Inc. As you read, make a flow chart that shows the production of the loonie. Keep in mind all the chemistry you have learned as you read.



[1]

Check your answers by turning to the Appendix, Section 2: Activity 3.

Nelson CHEMISTRY

Now that you have learned about various processes that use electrolysis, it is time to put your knowledge into practice by doing Investigation 13.7 on page 417 of *Nelson Chemistry*. This investigation will help to reinforce what you have learned.

Investigation 13.7: Copper Plating

Science Skills

- ☑ A. Initiating
- ☑ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☑ F. Evaluating

Perform Investigation 13.7: Copper Plating in order to evaluate the process and product of copper plating. Pay special attention to the required components, safety aspects, and applied science skills.

[1] Sherritt Inc., Fort Saskatchewan, Alberta is a Canadian public company based in Alberta. Sherritt's integrated operations include the production and marketing of fertilizers and chemicals, the production and sale of oil and natural gas, the refining and sale of nickel and cobalt, and the development and marketing of advanced industrial materials and metallurgical technologies.

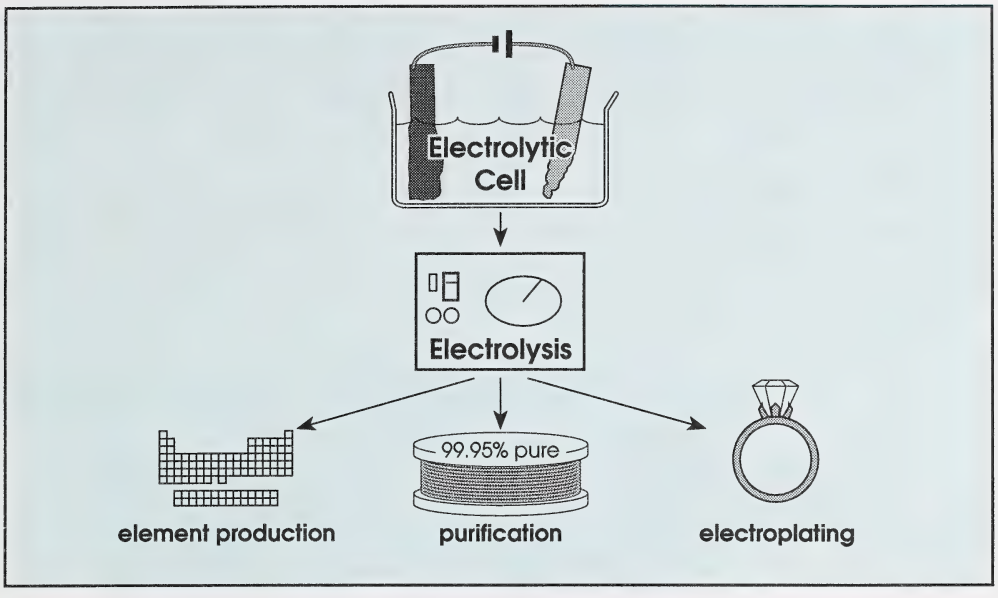
Guidelines

- Try to choose metal objects that will not react spontaneously with $Cu^{2+}_{(aq)}$.
- Once you have chosen your metal object, you should use that same type of object for all tests.
- Suggested ways of cleaning the objects to be plated may include soaking the object in an acid such as vinegar, cleaning it with a household cleaner before placing in solution, and cleaning it with steel wool. If you use a liquid cleaner, make sure you rinse the object before electroplating.

- You may wish to manipulate the concentration of the solution, or add a little acid to the solution before you electrolyse it. This may improve the plating process. Make sure you do not add too much acid.
- You may also wish to manipulate the voltage (up to 6 volts) in your power supply.
- Always start your electrolysis with the lowest voltage and slowly increase it if required.
- Use 75 mL of solution in a 100 mL beaker for each trial, or 20 mL of solution in a U-tube, depending on the objects you are plating.
- Obtain a minimum of six different results.

9. Make a prediction of the possible variables to be manipulated and controlled. Also predict the responding variable.
10. Make a list of your required materials.
11. Write a short procedure in step form.
12. Construct an appropriate evidence table, perform your procedure, and record your observations.
13. As your analysis, based on the evidence you obtained, what is the best process for electroplating copper?
14. As your evaluation, what improvements could you make to this investigation to increase the copper adhesion to your object?

Check your answers by turning to the Appendix, Section 2: Activity 3.

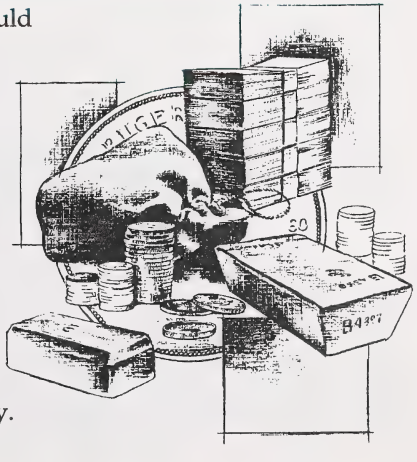


In this activity, you learned that even a product as simple as aluminum foil involves electrolysis. But electrical wiring, jewellery, chrome-plated bumpers, and even the coins in your pocket also involve converting electrical energy into chemical energy by the same process. Where else does electrolysis fit into your life?

ACTIVITY 4

Cell Stoichiometry

If you wanted to count a large quantity of coins, would you count them individually or would you stack them into piles of equal numbers and count the piles? And how would you measure how much water empties from a river into Lesser Slave Lake in one day? You certainly wouldn't use buckets! In this case, you might measure the distance the water in the river travels for a short period of time, determine the width and depth of the river, and use a formula. Obviously, your method of counting depends on what you are counting. So, how do you count the electrons used in a redox reaction? Knowing the amount of electrons flowing is important to the study of electrochemistry.



current – the rate at which electrons travel past a particular point; measured in amperes (A) – 1 A = 1 coulomb of charge passing a given point in 1 s

In this activity, you will find out how scientists use electric **current** to count electrons so they can determine the amount of reactant used or product gained in a chemical reaction. Since the electrons moving through a wire are what cause an electric current, by measuring the current, you are counting electrons. To understand how current can be used to count electrons, answer the next two questions.

1. a. From the introduction, you saw that it is possible to measure the volume of river water flowing into a lake in one day. Use this as an analogy to see how you can determine the number of electrons moving through a wire. Match the river characteristic with each of these electron characteristics: total charge transferred, current, time.

River	Electrons
volume of river water that passes one point in one minute	
how long you take to measure the river water volume; one day	
total water volume moved in the river in one day	

b. What, then, is the relationship between the three variables in the water analogy? in the electron relationship?

2. If one egg costs \$0.115 (11.5¢), you can determine the cost of one dozen eggs by multiplying \$0.115 × 12. Using correct units, your equation would look like this:

$$\frac{\$0.115}{1\ \cancel{egg}} \times \frac{12\ \cancel{eggs}}{1\ \cancel{dozen}} \times 1\ \cancel{dozen} = \$1.38 \text{ total}$$

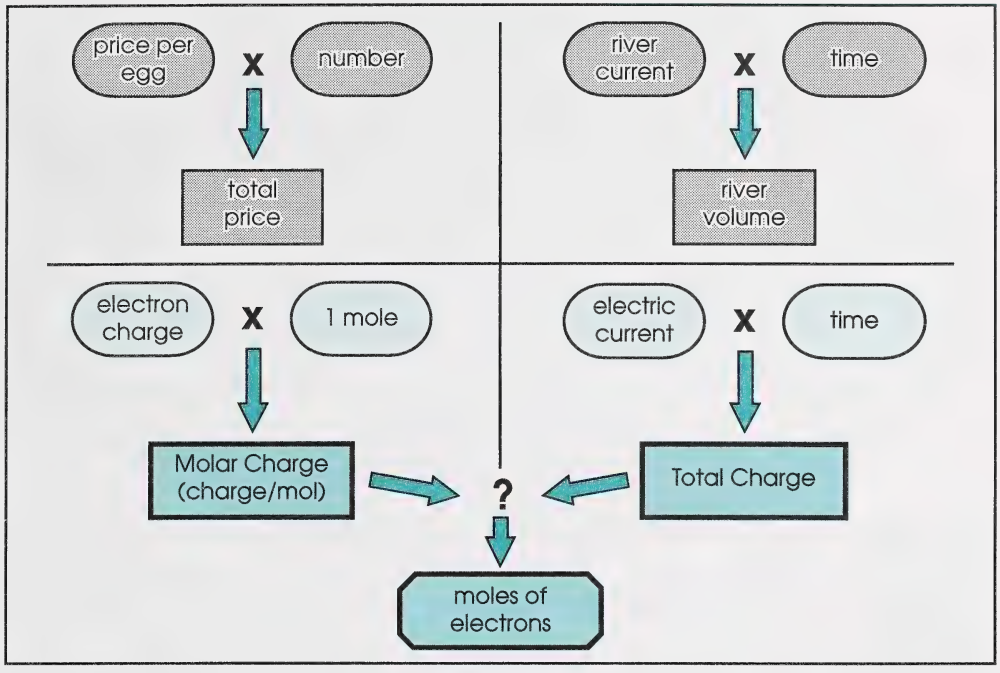
Use the same process to determine the charge on one mole of electrons. The charge on one electron is 1.6022×10^{-19} C and there are 6.02×10^{23} (Avogadro's number) electrons in one mole.

C – the symbol for the unit coulomb, the unit of electrical charge – 1 C = 6.24×10^{18} e^-

Check your answers by turning to the Appendix, Section 2: Activity 4.

Faraday constant (F) – the electrical charge on one mole of electrons; 96 500 C/mol

In the previous questions, you saw that you could determine the total charge transferred through a wire if you knew the current and the time, and you also determined the charge of one mole of electrons. This molar charge of electrons is called the **Faraday constant (F)**. How can you use this information to calculate the number of moles of electrons $\left(n_{e^-}\right)$ transferred in a wire, and how will this help you determine mass changes of the electrodes in an electric cell?



In fact, calculating moles of electrons is very similar to calculating moles using molar mass or moles using molar heats. Do the next question to confirm the similarity.

3. Describe how to determine the following:

   a. moles of a compound given its mass and molar mass

   b. moles of a compound given its heat and molar heat

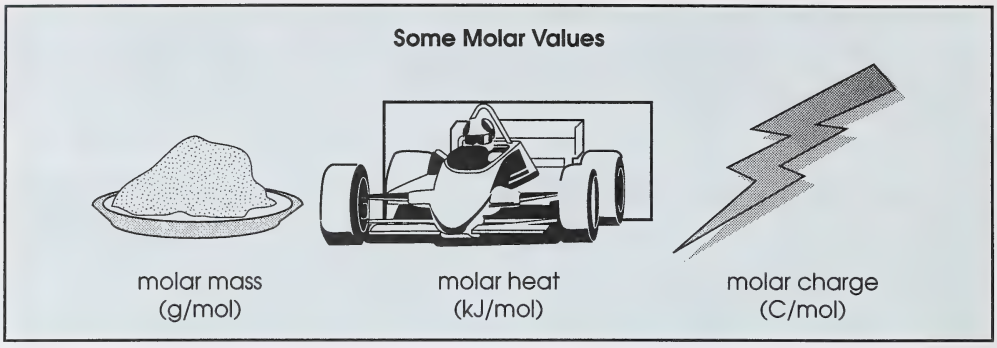
   c. moles of electrons given their charge and molar charge

4. Calculate the number of moles of electrons $\left(n_{e^-}\right)$ for the following:

   a. The charge transferred in a wire is 1.93×10^4 C and the molar charge is 9.65×10^4 C/mol.

   b. The charge gained by the cathode is 722 C and the charge on one mole of electrons is 9.65×10^4 C/mol.

Check your answers by turning to the Appendix, Section 2: Activity 4.



As you can see, calculating the moles of electrons gained by the cathode or lost from the anode is very similar to calculating the mass from molar mass or heat from molar heat. In a chemical reaction, then, you might expect that the stoichiometric calculations involving electrons would be similar, and you would be right.

Nelson CHEMISTRY

Read Section 13.4: Stoichiometry of Cell Reactions on pages 417-418 of your textbook; then answer the next three questions to check your understanding of cell stoichiometry and practise some examples.

5. Fill in the blanks with the appropriate words or terms.

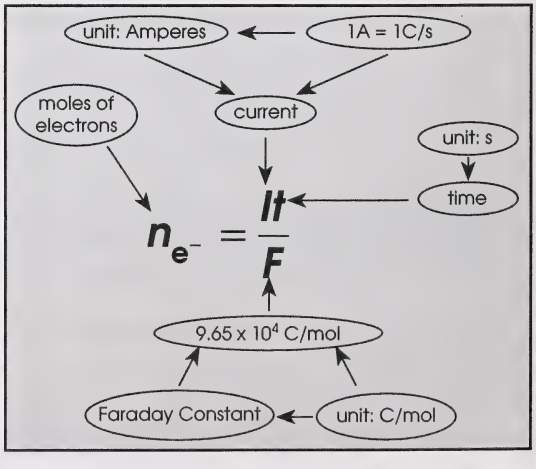
   Another name for the charge on one mole of electrons is the _____ constant. The letter used to represent this constant is _____. It has units of _____. Likewise, current is measured in _____ and is represented by the letter _____. One ampere is equivalent to one _____ per second. The coulomb is a unit of charge that is represented by the letter _____. When dealing with current, time is always measured in _____.

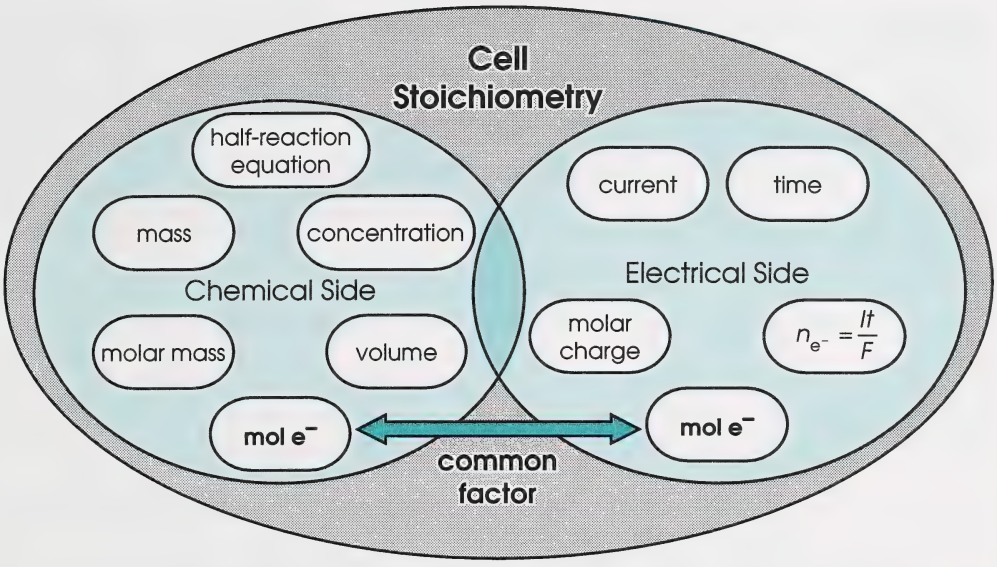
6. Determine the number of moles of electrons moved in each of the following:

   a. A current of 6.2 A flows for 40 s.

   b. A current of 1.44 A flows for 50 min.

   c. A current of 1.00 kA flows for 7.0 h.



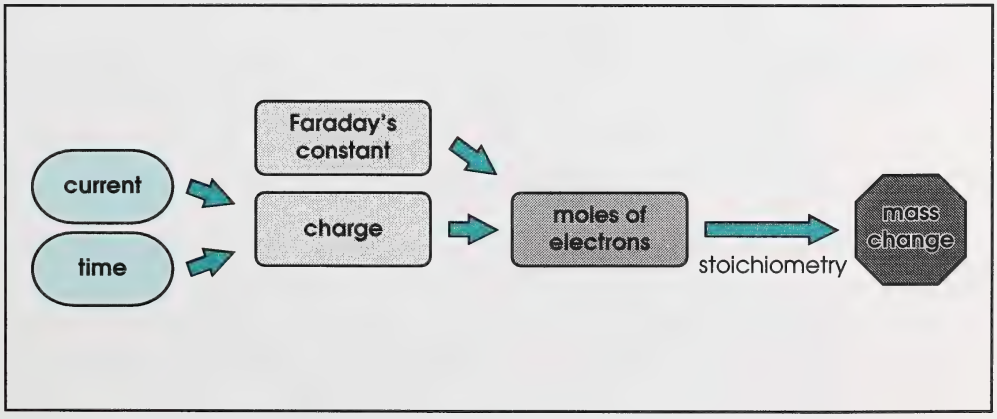


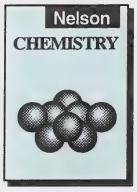
Nelson CHEMISTRY

7. Answer textbook questions 34, 35, 36, and 37 on pages 418-419 of *Nelson Chemistry*.

Check your answers by turning to the Appendix, Section 2: Activity 4.

Think back to Investigation 13.7: Copper Plating in the previous activity. When you increased the voltage, the current increased and more copper was plated onto your cathode. Likewise, if you let the current run longer, you got a thicker copper layer. From your stoichiometric calculations you can now figure out why. The process is summarized in the following diagram.





In fact, if you had measured the current and the time, you could have predicted how much copper was collected at the cathode. Lab Exercise 13D: Quantitative Electrolysis on page 419 of *Nelson Chemistry* will give you the opportunity to see how these types of calculations work in the laboratory environment.

Lab Exercise 13D: Quantitative Electrolysis

Science Skills

- ☑ A. Initiating
- ☐ B. Collecting
- ☐ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☑ F. Evaluating

Complete Lab Exercise 13D, paying special attention to the applied science skills. The following three questions lead you through the lab exercise.

8. Predict the mass of tin that should be plated at the cathode.

9. What mass of tin was actually produced in the experiment?

10. a. Determine the percentage error in your results.

    b. Discuss whether this amount of error is reasonable and state two possible reasons for it.

Check your answers by turning to the Appendix, Section 2: Activity 4.

So far, you have solved cell stoichiometry calculations by being given the current and time. Do you think these calculations could be done without knowing the current or the time? Study the following example to see one type of question where this can be done.

Example

If 10.792 g of $Al_{(s)}$ were produced at the cathode of a cell in an electrolysis experiment, what mass of $Zn_{(s)}$ would be produced in a different electrolytic cell using the same current and same time?

Solution

If the same current and same time were used, then the same number of electrons must have been transferred since the same charge was transferred. Therefore, balance the electrons in the two half-reactions and do your stoichiometry as usual.

Aluminum Half-Reaction

$2\left(Al^{3+}_{(aq)} + 3\,e^- \rightarrow Al_{(s)}\right)$

$2\,Al^{3+}_{(aq)} + 6\,e^- \rightarrow 2\,Al_{(s)}$

10.792 g

26.98 g/mol

Zinc Half-Reaction

$3\left(Zn^{2+}_{(aq)} + 2\,e^- \rightarrow Zn_{(s)}\right)$

$3\,Zn^{2+}_{(aq)} + 6\,e^- \rightarrow 3\,Zn_{(s)}$

65.38 g/mol

$m = ?$

$$m_{(Zn)} = 10.792\ \cancel{g\ Al} \times \frac{1\ \cancel{mol\ Al}}{26.98\ \cancel{g\ Al}} \times \frac{3\ \cancel{mol\ Zn}}{2\ \cancel{mol\ Al}} \times \frac{65.38\ g}{1\ \cancel{mol\ Zn}}$$

$$= 39.23\ g$$

The mass of Zn produced would be 39.23 g.

As you can see, you are not always asked to find the same information in every electrolysis calculation. Do the next three questions to practise solving for the different variables found in these types of problems.

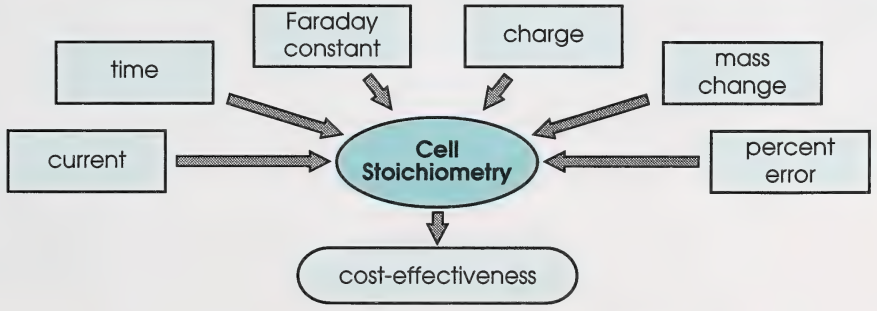
11. If 0.300 g of gold were plated onto a necklace from a solution of $Au(NO_3)_{3(aq)}$, what mass of silver could be plated onto a similar necklace using the same current and same time from a $AgNO_{3(aq)}$ solution?

12. An identical current is run through three electrolytic cells for the same length of time. The three cells contained $Zn(NO_3)_{2(aq)}$, $Ni(NO_3)_{3(aq)}$, and $Pb(NO_3)_{4(aq)}$. Which of these cells will create the greatest mass increase at the cathode?

Nelson CHEMISTRY

13. Do textbook questions 24 and 25 on page 422 of *Nelson Chemistry*.

Check your answers by turning to the Appendix, Section 2: Activity 4.



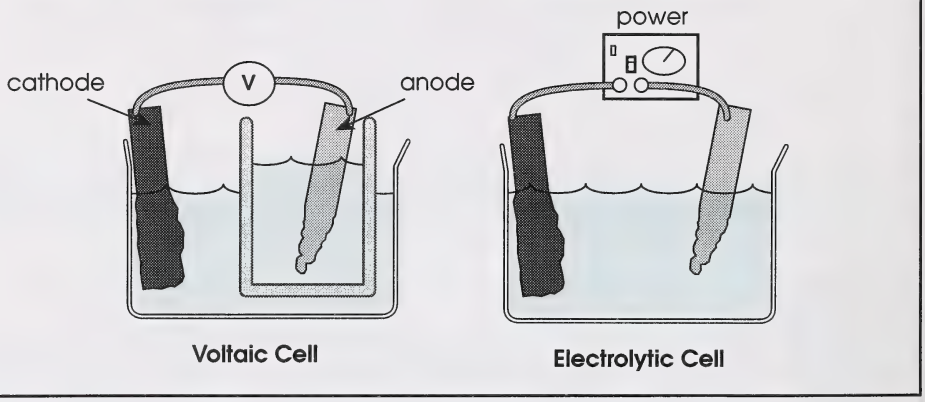
In today's business world, everyone must keep track of expenses and profits, or losses and gains. The same thing is true for industries that use electrolysis. In order to do that, the manufacturers have to maintain a count of the number of electrons that are being consumed. They do this by measuring the current and time. In addition, they keep track of the amount of mass produced at the cathode and anode in order to know how efficiently the electrolysis process is working. In this activity, you have learned how to do some of these calculations, as well as learning some of the theory behind them.

Follow-up Activities

If you had difficulties understanding the concepts in the activities, it is recommended that you do the Extra Help. If you had a clear understanding of the concepts, it is recommended that you do the Enrichment.

Extra Help

1. The electrolytic cell is very similar to the electrochemical or voltaic cell because both cells involve two electrodes and an electrolyte, and in both cells the electron transfer occurs through an external circuit.



For each of the following statements, decide whether the statement is true for the electrolytic cell (write E), voltaic cell (write V), both cells (write B), or neither cell (write N).

a. Reduction occurs at the cathode.

b. Oxidation occurs at the anode.

c. Anions move to the anode.

d. Cations move to the cathode.

e. Electrons move from the anode to the cathode.

f. A salt bridge, porous cup, or other semi-permeable membrane is needed.

g. This can be a power source for electrolysis.

h. ΔE° is positive.

i. The oxidizing agent is lower than the reducing agent on the table of redox half-reactions.

j. The reaction is non-spontaneous.

k. A power source is required.

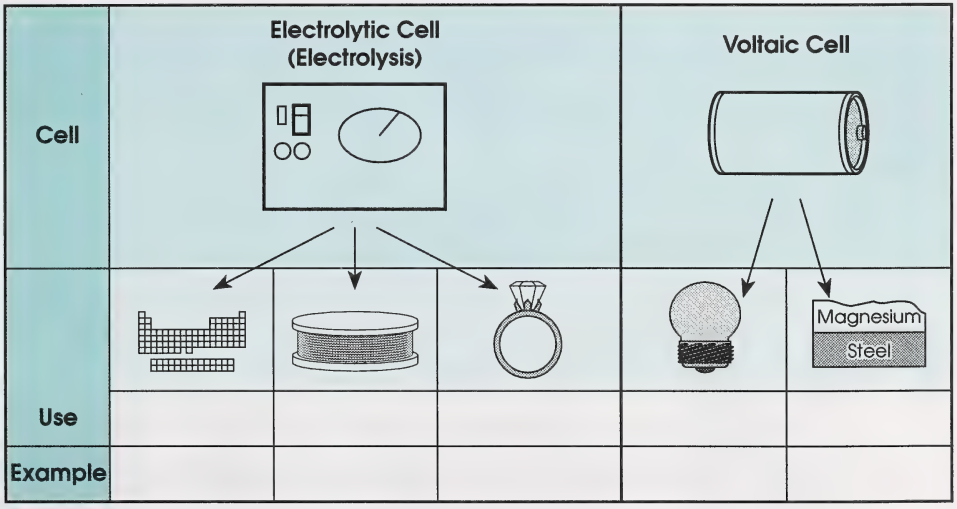
l. $\Delta E^\circ = E^\circ_r(\text{cathode}) - E^\circ_r(\text{anode})$

m. The cathode usually gains mass.

n. The anode can lose mass.

o. It converts electrical energy to chemical energy.

2. The voltaic cell converts chemical energy into electrical energy and has many special uses. Because the electrolytic cell converts electrical energy to chemical energy, it is used in different situations. As indicated in the following diagram, write the three main uses for an electrolytic cell and the two main uses of a voltaic cell; then give an example of each use.



3. When calculating the mass gained at the cathode or lost at the anode in an electrolytic cell, you need to know the substances involved, the current, and the time. The following steps are the ones that you need to follow in order to determine the mass changes that occur at the electrodes. Arrange them in their proper order.

a. Multiply the moles of electrons by the factor *number of mol element/number of mol* e^-. This uses the chemical equation to get the number of moles of the element and electrons.

b. Multiply current and time to determine the total charge.

c. Convert time to seconds.

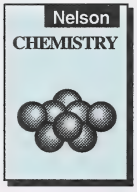
d. Multiply by the molar mass of the element to get the mass produced.

e. Divide the total charge by Faraday's constant $\left(9.65 \times 10^4 \text{ C/mol}\right)$ to get moles of electrons.

4. Predict the mass gained by the cathode when a current of 4.5 A runs through a solution of $ZnBr_{2(aq)}$ for 20 minutes.

5. Predict the mass of product formed at the anode of a molten electrolytic cell when a current of 50.0 A is used to electrolyse liquid $MgO_{(l)}$ for 1.5 hours.

Check your answers by turning to the Appendix, Section 2: Extra Help.



If you require further questions for practice, you can answer any of the appropriate questions in the Review and Applications sections of the Overview on pages 420-422 of *Nelson Chemistry*. The answers to these questions can be found on pages 516-518 of your textbook.

Enrichment

Complete **one** of the following.

1. Interview someone involved in refining metals from ores. Compare the processes and costs of refining aluminum from its ore versus recycling it. Alternately, you could do some library research into this topic.

2. If you have the opportunity, make an appointment to talk to someone at the Sherritt Inc. plant in Fort Saskatchewan to compare the processes used to make the loonie with the processes used to make other coins such as the nickel.

PHOTO COURTESY GOVERNMENT OF SASKATCHEWAN

Science Skills

- ☑ A. Initiating
- ☑ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☑ E. Synthesizing
- ☑ F. Evaluating

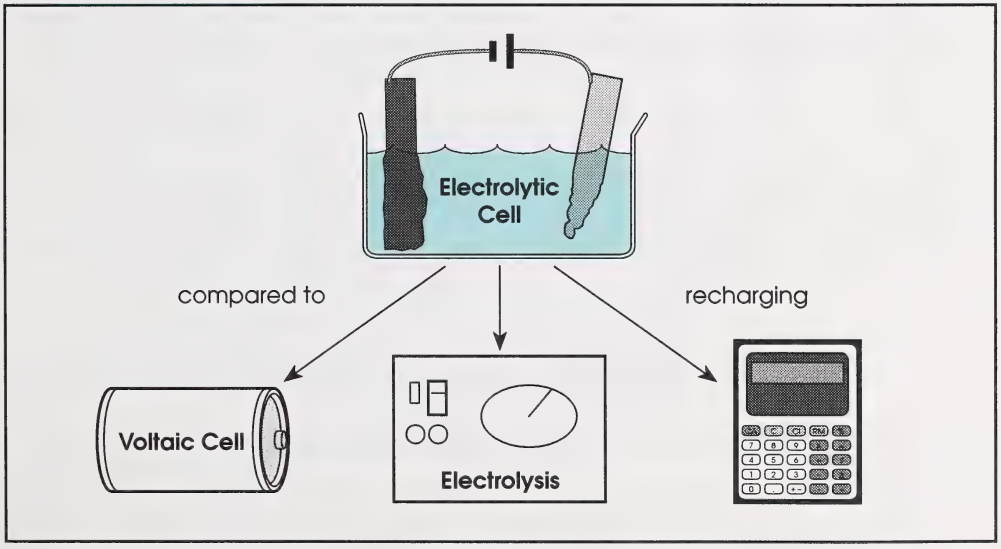
3. In all electric cells, chemical energy is converted to electrical energy or electrical energy is converted to chemical energy. To measure the energy, you can use the equation $E = VIt$, where E stands for energy in joules, V is voltage, I is current, and t is time in seconds. If you cannot measure the current, you can measure the resistance of the circuit instead, and use the equation $E = V^2t / R$ to find the energy output. R stands for resistance and is measured in ohms.

   Design an experiment that would estimate the amount of energy a battery will produce. Compare this to the mass of natural gas that would be required to produce the same amount of energy.

   Pay special attention to the required components, safety aspects, and applied science skills.

Check your answers by turning to the Appendix, Section 2: Enrichment.

Conclusion



In this section, you studied the electrolytic cell and compared it with the voltaic cell. By electrolysing various solutions, you learned about the basic principles of electrolysis and some of the ways in which electrolysis is used, such as making coins and making chemicals. You also learned how to carry out chemical calculations related to electrolysis.

40

Section 2 Assignment: Turning Electricity into Chemistry

Review the Evaluation information found in the introductory pages of this module.

It is important to number and clearly identify each page with the following information at the top:

Chemistry 30 -- Module 4 Section 2 Assignment Page # Name and ID#

Be sure to write legibly. Leave a wide left margin and number all of your pages.

1. Decide whether each of the following statements concerning an electrolytic cell are true (T) or false (F). For each statement, explain why it is true or false. **(8 marks)**

   a. A salt bridge or other semi-permeable membrane is required.

   b. The oxidation number of the species reacting at the anode increases as the cell operates.

   c. The reduction half-reaction has a higher reduction potential than the reduction potential of the oxidation half-reaction.

   d. The reduction potential of $Cd^{2+}_{(l)} + 2\,e^- \rightarrow Cd_{(l)}$ is -0.40 V during the electrolysis of molten cadmium chloride $\left(CdCl_{2(l)}\right)$.

   e. In the electrolysis of $NaBr_{(aq)}$, the solution will turn darker at the anode, while the mass of the cathode will increase.

   f. Tripling the current and halving the time in which a cell operates will cause two-thirds of the mass of the product to form at the cathode.

   g. A greater mass increase will occur at the cathode from a $CoCl_{2(aq)}$ solution than from a $NiCl_{3(aq)}$ solution with the same current over the same time.

   h. An electrochemical cell with a potential of 3.0 V can be used to produce hydrogen at the cathode in the electrolysis of aqueous lithium hydroxide.

Science Skills

- ☑ A. Initiating
- ☐ B. Collecting
- ☐ C. Organizing
- ☑ D. Analysing
- ☑ E. Synthesizing
- ☐ F. Evaluating

2. A student wanted to determine the purity of a sample piece of iron. The following equipment was available:

- the impure iron sample
- a piece of copper
- 1.0 mol/L $Zn(NO_3)_{2(aq)}$
- (2) 250 mL beakers
- a porous cup
- an electronic balance
- a piece of zinc
- 1.0 mol/L $Fe(NO_3)_{2(aq)}$
- 1.0 mol/L $Cu(NO_3)_{2(aq)}$
- steel wool
- wire
- voltmeter

The student constructed a cell using some of this material and recorded the following results:

	Electrode A	Electrode B
Initial mass (g ± 0.001 g) Final mass (g ± 0.001 g)	5.884 6.213	6.917 6.587
Mass change (g ± 0.002 g)	0.329	0.330

(Assume that the impurities are unreactive.)

a. Write the anode equation, cathode equation, and net ionic equation that represent the reaction that occurs. **(3 marks)**

b. Calculate the net potential difference of this cell. **(1 mark)**

c. Draw a sketch of the cell operating in this question. Label the cathode, anode, direction of cation and electron flow, and all materials used in the construction of the cell. You may use only the materials listed in this question. **(6 marks)**

d. Is the cathode Electrode A or Electrode B? Explain. **(2 marks)**

e. Determine the percent of iron in the impure sample. **(4 marks)**

f. State two other qualitative observations you would observe as this cell operates. **(2 marks)**

Response Page Setup

Reproduce the science skills box after your response to this question. Remember to indicate your evaluation of your skill level for each identified science skill. Leave a space for the teacher assessment.

	A.	B.	C.	D.	E.	F.
Self:	☐	☐	☐	☐	☐	☐
Teacher:	☐	☐	☐	☐	☐	☐

Response Page Setup

3. Calculate the net cell potential of the following cells, making sure you include the labelled cathode and anode half-reaction equations and the net ionic reaction equation in your answer. Indicate, as well, whether the reaction is spontaneous or non-spontaneous. Show all your work clearly. Marks will be given for both your work and your answer.

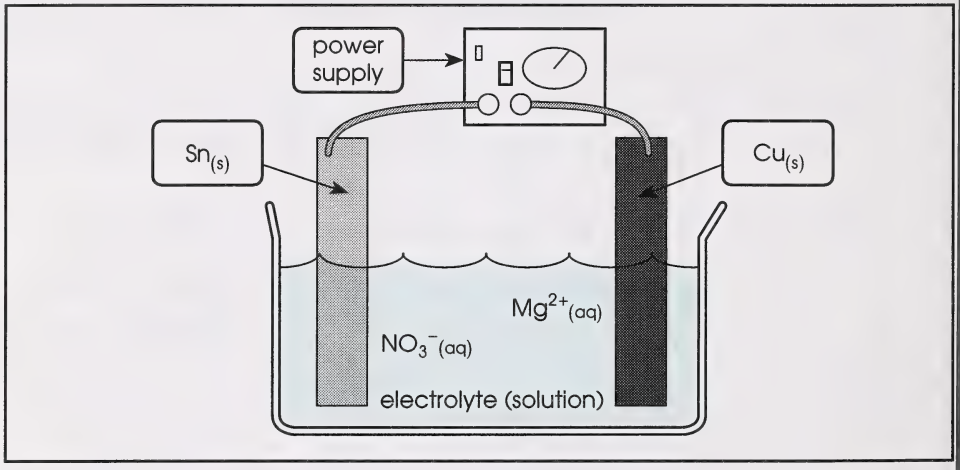
   a. A platinum and a copper electrode are immersed into a lead(II) nitrate solution and connected to make a cell. **(3 marks)**

   b. Two silver electrodes are placed in a solution of $CaI_{2(aq)}$ and connected to make a cell. **(3 marks)**

Response Page Setup

4. Calculate the mass of element produced at the anode during the electrolysis of molten chromium(III) bromide if 2.50 A runs through the cell for 45 minutes. Show all your work clearly. Marks will be given for both your work and your answer. **(3 marks)**

5. Use the following diagram to answer this question.

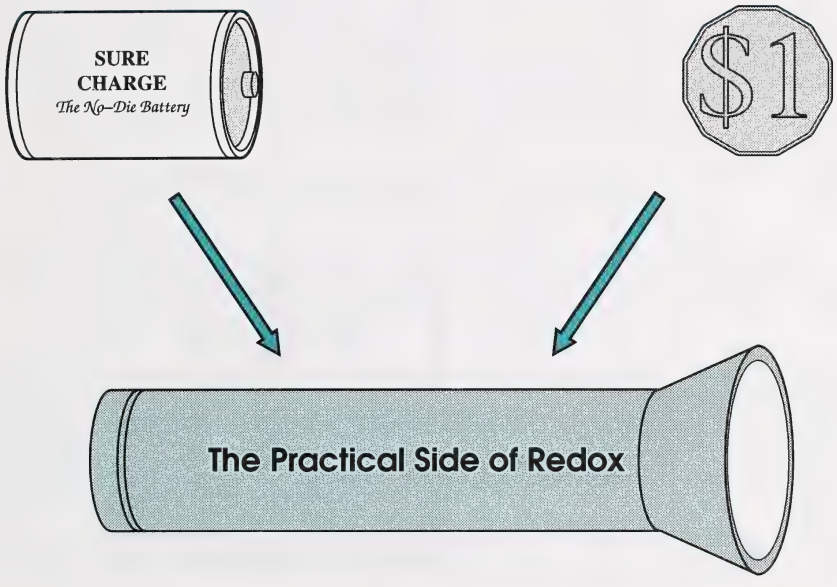


   a. Predict the strongest oxidizing and reducing agents in the given cell. Based on your answer, describe the necessary direction of electron flow in the cell (either from $Sn_{(s)}$ to $Cu_{(s)}$ or from $Cu_{(s)}$ to $Sn_{(s)}$). **(2 marks)**

Response Page Setup

   b. Calculate the time in hours needed to make the anode change in mass by 3.00 g if the current through the solution is 275 mA. Remember, 1 mA = 10^{-3} A. Show all your work clearly. Marks will be given for both your work and your answer. **(3 marks)**

MODULE SUMMARY



Throughout this module, you examined a variety of items – from the battery to the loonie. In the process, you discovered the large role that the basic principles of electric cells play in your life, in industry, and in society. In particular, you studied the major electrochemical processes involved in voltaic cells, ranging from the common battery to cathodic protection. You also learned how voltaic cells were constructed. Later, you compared the voltaic cell with the electrolytic cell and saw how the principles of the electrolytic cell were applied to the production of the loonie, as well as to metal refining and element production. Finally, you practised various calculations involved in electrochemical processes.

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Final Module Assignment

Review the Evaluation information found in the introductory pages of this module.

It is important to number and clearly identify each page with the following information at the top:

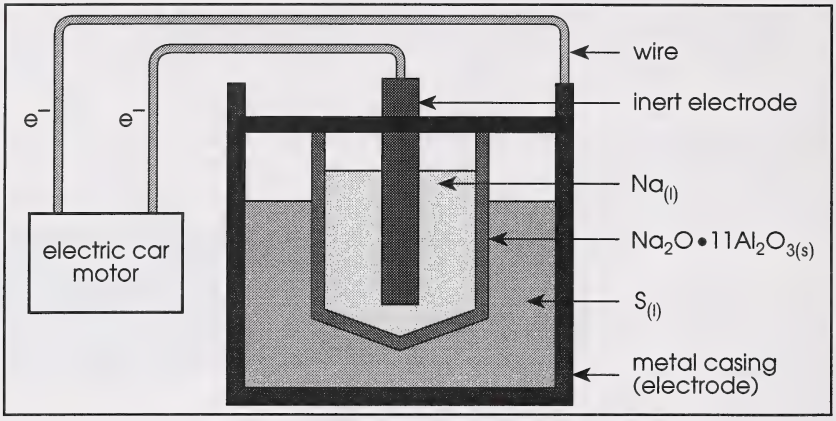
Chemistry 30 – Module 4 Final Module Assignment Page # Name and ID#

Be sure to write legibly. Leave a wide left margin and number all of your pages.

Science Skills

- ☐ A. Initiating
- ☐ B. Collecting
- ☐ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☐ F. Evaluating

1. An alternative to the lead-acid batteries used in electric vehicles is the sodium-sulphur battery. In the sodium-sulphur battery, sodium is reacted according to the half-reaction $Na_{(l)} \rightarrow Na^{+}_{(l)} + e^{-}$, and sulphur is reacted according to the half-reaction $3\,S_{(l)} + 2\,e^{-} \rightarrow S_3^{2-}{}_{(l)}$. The reaction is run at high temperature; the materials used in the construction of the cell are molten sodium, molten sulphur, and a solid electrolyte called *beta alumina* with the formula $Na_2O \bullet 11Al_2O_3$.



Response Page Setup

Copy the preceding diagram (yours should be similar in size) onto your response page. Indicate the cathode, anode, and porous membrane beside the appropriate label on the diagram. Indicate the direction of electron flow using arrows placed above or below the e^{-} symbols on the diagram. **(4 marks)**

Reproduce the science skills assessment box after your response to this question. Remember to indicate your evaluation of your skill level for each identified science skill. Leave a space for the teacher assessment.

Self:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐
Teacher:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐

Science Skills

- ☐ A. Initiating
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- ☑ F. Evaluating

2. Use the following information to answer this question.

The lead in a lead-acid battery oxidizes according to the half-reaction equation $Pb_{(s)} + SO_4^{2-}{}_{(aq)} \rightarrow PbSO_{4(s)} + 2\,e^{-}$. The oxidation half-reaction in the sodium-sulphur battery is represented by $Na_{(l)} \rightarrow Na^{+}_{(l)} + 1\,e^{-}$, and the oxidation half-reaction in a lithium-polymer battery, currently under development, is represented by $Li_{(s)} \rightarrow Li^{+}_{(s)} + e^{-}$.

a. If 500 g of lead are oxidized in a lead-acid battery, determine the mass of sodium and lithium that would be oxidized by the same current for the same time in the other batteries. **(3 marks)**

b. Based on your answer to part a., which battery do you think would be best to use for electric automobiles from a technological perspective? Explain, considering such factors as space and energy density. **(2 marks)**

Response Page Setup

Show your work clearly for each part. Marks will be given for both your work and your answer.

Reproduce the science skills assessment box after your response to this question. Remember to indicate your evaluation of your skill level for each identified science skill. Leave a space for the teacher assessment.

Self:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐
Teacher:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐

Science Skills

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- ☑ E. Synthesizing
- ☐ F. Evaluating

3. One of the problems with electric vehicles is the small amount of energy a battery can produce per kilogram compared with gasoline.

a. A mass of 1.00 kg of gasoline will generally produce 43.7 MJ of energy. If one particular car can travel 700 km on one tank (50 L or 37 kg) of gasoline, how much energy would the car produce from one tank of gasoline? **(2 marks)**

b. If only 20% of the energy produced by burning the gasoline is used to move the car forward, how much energy does it take to move the car 700 km? **(2 marks)**

c. A typical 15 kg lead-acid car battery will only produce 1.3×10^2 kJ/kg, but it operates at 90% efficiency. How many lead-acid batteries would be required to make the same car travel the same distance as in 3.b.? **(3 marks)**

Response Page Setup

Show your work clearly for each part. Marks will be given for your work as well as your answer.

Reproduce the science skills assessment box after your response to this question. Remember to indicate your evaluation of your skill level for each identified science skill. Leave a space for the teacher assessment.

Self:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐
Teacher:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐

Science Skills

- ☐ A. Initiating
- ☐ B. Collecting
- ☐ C. Organizing
- ☑ D. Analysing
- ☑ E. Synthesizing
- ☑ F. Evaluating

4. An electrolytic cell was made of two carbon electrodes immersed in a 1.0 mol/L nickel(II) iodide solution.

   a. Write the net ionic equation for the reaction that occurs during electrolysis. **(1 mark)**

   b. Calculate the net cell potential. **(1 mark)**

   c. Draw a potential energy diagram for this reaction (as you did in Modules 1 and 2). Remember to label your axes as the Reaction Progress and E_p (kJ), and indicate the relative positions of the reactants and products. **(2 marks)**

Response Page Setup

Show your work clearly for each part.

Reproduce the science skills assessment box after your response to this question. Remember to indicate your evaluation of your skill level for each identified science skill. Leave a space for the teacher assessment.

Self:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐
Teacher:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐

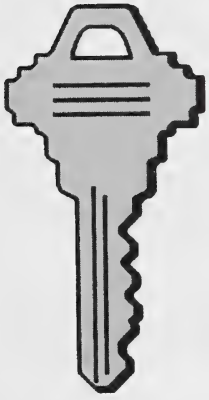
To ensure that all of your work has been completed in a satisfactory manner, check off the items in the following list:

- ☐ Section 1 Assignment has been completed.
- ☐ Section 2 Assignment has been completed.
- ☐ Final Module Assignment has been completed.
- ☐ Your responses are organized and neat, with room for teacher comments.
- ☐ All your response pages are numbered consecutively and identified with this heading.

Chemistry 30 – Module 4 Section # Assignment Page # Name and ID #

Submit **only** your **assignment response pages** for evaluation.

Appendix



Glossary
Suggested Answers

Glossary

ammeter: a device used to measure electrical current

anomaly: an unexpected occurrence

batteries: two or more voltaic cells connected in series, or one or more voltaic cells in a finished package

C: the symbol for the coulomb, the unit of electrical charge – $1\ C = 6.24 \times 10^{18}\ e^-$

cathodic protection: the method of forcing a metal to become a cathode by attaching a stronger reducing agent to it

charge: a quantity of electricity; is measured in coulombs (C)

corrosion: the oxidation of any material, especially metals

current: the rate at which electrons travel past a particular point; measured in amperes (A) – 1 A = 1 coulomb of charge passing a given point in 1 s

electrochemical cell: any electrochemical reaction in which an external current is made from a chemical reaction or a chemical reaction is made from an electric current

electrodes: solid material through which electricity flows; usual site of electrical connections

electrode potential: another name for reduction potential

electrolysis: the process of forcing a non-spontaneous redox reaction to occur using electrical energy

electrolyte: a substance that, when dissolved in water, conducts an electric current

electrolytic cell: a non-spontaneous cell that converts electrical energy into chemical energy

electroplating: covering a material with a thin layer of protective metal by the process of electrolysis

Faraday constant (F): the electrical charge on one mole of electrons; 96 500 C/mol

multimeter: a device that can be used to measure a variety of electrical quantities such as current, voltage, and resistance

potential difference: the difference in the amount of potential energy per unit charge between any two points; measured in volts (V); 1 volt = 1 coulomb of charge having 1 J of energy

power: the rate at which energy is being consumed in watts (W); 1 watt = 1 joule of energy being used in 1 s

sacrificial anode: a metal that is oxidized (a stronger reducing agent) to protect another metal from corrosion

series: the positive electrode of one cell hooked up to the negative electrode of another cell so that the cell voltages are additive

voltaic cells: any cells that produce an electric current from a spontaneous electrochemical reaction

voltmeter: a device used to measure the potential difference between any two points in a circuit

Suggested Answers

Section 1: Activity 1

1. Depending on the hardness of your water, the conductivity of the tap water should be approximately 0.1-0.2 μA, while the conductivity of the salt solution should be about 0.4-0.6 μA.

2. Your results should be similar to those listed in the following table. Factors such as how juicy the fruit is will give a wide range of results.

Trial	Metal 1	Metal 2	Material Between Metals	Ammeter Reading (mA)	Voltmeter Reading (Volts)
1	copper	zinc	salt water/towel	3.60	0.72
2	copper	lead	salt water/towel	0.45	0.39
3	copper	Mg	salt water/towel	9.20	1.27
4	zinc	lead	salt water/towel	1.85	0.33
5	zinc	Mg	salt water/towel	7.10	0.66
6	lead	Mg	salt water/towel	5.90	0.94
7	copper	zinc	orange	0.61	0.88
8	copper	lead	orange	0.23	0.50
9	copper	Mg	orange	0.63	1.78
10	zinc	lead	orange	0.20	0.39
11	zinc	Mg	orange	0.14	0.97
12	lead	Mg	orange	0.50	1.38
13	copper	zinc	potato	0.17	0.65
14	copper	lead	potato	0.08	0.37
15	copper	Mg	potato	0.32	1.66
16	zinc	lead	potato	0.10	0.38
17	zinc	Mg	potato	0.17	0.93
18	lead	Mg	potato	0.27	1.29

3. The immediate evidence that a chemical reaction was occurring should be that a current and a potential difference (voltage) were measurable.

4. Galvani's hypothesis was incorrect because whenever any two different metals contact an electrolytic solution a current will flow. Galvani's prediction was simply too specific since muscle tissue is only one substance that contains electrolytes.

5. It was necessary to keep the two metals from touching so that the electrons would move through the wire rather than directly from one metal to another, since transferring by direct contact is easier than by transferring through a wire.

6. The material between the two metals must be conductive (electrolytic) so that a complete circuit is created.

7. a. Electrochemical cell, voltaic cell, and galvanic cell generally mean the same thing.

   b. All of the examples would be analogous to a battery except the doughnut because a battery is made up of a series of cells or units arranged in a continuous line. The doughnut would be analogous to a single cell (a voltaic cell).

8. The following is one possible solution. All the electrolyte materials are used.

Object to Operate	Voltage Required	Amount and Type of Electrolyte Required	Voltage Generated
car battery	12 V	14 lemons x 0.86 V	12 V
camera flash	6 V	7 lemons x 0.86 V	6.0 V
transistor radio	9 V	45 apples x 0.20 V	9.0 V
flashlight	3 V	15 apples x 0.20 V	3.0 V
ghetto-blaster	9 V	60 potatoes x 0.15 V	9.0 V

9. **Textbook question 1:**

The three main parts of any simple electric cell are the two electrodes (anode and cathode) made of different conducting materials and an electrolytic solution separating them.

Textbook question 2:

Empirically, an electrode is any solid material that conducts electricity; an electrolyte is any solution that conducts electricity.

Conventionally, an anode is the negative electrode; the cathode is the positive electrode.

Textbook question 4:

From a scientific perspective, when batteries are placed in backwards to the diagram on the device, the electrons flow in the wrong direction. Correct placement ensures that electrons flow in the correct direction.

From a technological perspective, the diagram shows how the batteries should be placed to make the appliance work properly and not damage it.

Textbook question 5:

Electric current is a measure of the rate of movement of electrons past a given point; voltage is a measure of the energy difference per unit electric charge.

10. a. potential difference
b. charge
c. increase
d. remain the same
e. current
f. anode
g. cathode
h. anode
i. cathode
j. power
k. energy density
l. charge
m. total energy

11. You experience a shock if you have metallic-based fillings and chew on aluminum foil because you have all the requirements for an electric cell. Since the metals used in your fillings are not aluminum, you have two electrodes of different metals – the foil and your fillings. Your saliva is the electrolyte. When the electrons begin to flow from the aluminum to the filling, a small current is generated, passing through the nerves in your teeth creating a sensation of electric shock.

Section 1: Activity 2

1. The best way to be systematic when trying to improve upon an experimental design is to manipulate only one variable at a time, keeping the other variables controlled.

2. Your prediction should be similar to the following:

   From Investigation 13.1, it would be reasonable to predict that with the materials available, the copper wire would form the best cathode for the aluminum-can cell because whenever copper was used in Investigation 13.1, it most consistently created the greatest voltage. Likewise, the electrolyte that gave the greatest voltage was the vinegar. Therefore, the vinegar would probably be the best electrolyte. If you rely on your knowledge from Module 3, you might predict that since aluminum and copper are far apart on the redox table, their cell might produce the highest voltage.

3. Your procedure should be similar to the following:

   - Remove the top of the aluminum can with a can-opener or tin snips.
   - Polish all the metals with steel wool.
   - Make sure the graphite core of the pencil is bare at both ends.
   - Fill the aluminum can 2/3 full with salt water.
   - Use a multimeter to test the voltage of each metal in the salt-water cell.
   - Discard the salt solution.
   - Rinse the can.
   - Replace the salt solution with vinegar and retest with each metal.
   - Discard the vinegar.
   - Rinse the can.
   - Replace the vinegar with a baking soda solution and retest with each metal.

4. Your evidence table should look similar to the following:

VOLTAGE READINGS FOR ALUMINUM-CAN CELL

Electrodes / Electrolytes	Copper Wire	Lead Weight	Pencil Lead	Galvanized Metal	Iron Nail
Salt Solution	0.44 V	0.16 V	0.63 V	0.26 V	0.24 V
Vinegar	0.47 V	0.08 V	0.55 V	0.43 V	0.08 V
Baking Soda	0.41 V	0.04 V	0.40 V	0.47 V	0.29 V

5. a. The best electrolyte would be salt water and the best electrode would be a pencil lead.

   b. Your actual decision is not as important as the reasons you give. Your answer should take into account such things as the following:

      - Lead is environmentally unfriendly.
      - Carbon is cheapest and readily available.
      - Copper is the most expensive per unit mass.
      - Salt solution is readily available.

   In other words, the electrolyte and electrode chosen should have the greatest availability, least cost, and minimum impact on the environment.

   c. The technological developments pursued depend on a balance between the costs involved in production, the demands by society, the availability of the resources, and the impact on the environment.

6. Some examples for improvements may include

   - increasing the concentration of the electrolyte
   - replacing the copper with a metal that would work even better
   - inserting a soaking paper towel into the cell to help prevent spillage, yet maintaining the electrolyte

7. a. primary cell
   b. fuel cell
   c. primary cell
   d. secondary cell
   e. fuel cell
   f. primary cell

8. Carbon cells are dry cells that have a short life span. Alkaline cells, because of their construction, have a much longer life. This is independent of the brand of battery. Therefore, Omar's conclusion should have been that alkaline batteries last longer. Stacey likely compared two alkaline batteries, and found that the SURE-CHARGE alkaline was a better alkaline design.

9. **Textbook question 6:**

   Scientific knowledge helps you explain how and why a device works. Technological problem solving helps you develop a device or improve on the design of a device.

   Textbook question 10:

   The unique design of the Molicel® is the fact that the cathode and anode are rolled sheets of metal layered with electrolyte, rather than distinct parts of the cell being the anode and cathode. This allows a greater current to flow.

   Textbook question 11:

   The technological problem solved by a fuel cell is that you can add more reactant as it is used up. This considerably increases the life of the cell compared to a regular dry cell.

10. In making your decision, you should examine the advantages and disadvantages of each. The alkaline cell would probably be best because of its long initial life. Rechargeables are more suited to items that use a lot of power and would normally require a frequent cell change anyway.

ALKALINE CELL		RECHARGEABLE CELL	
Advantages	**Disadvantages**	**Advantages**	**Disadvantages**
• longer life per discharge • lower initial cost	• primary cell • less environmentally friendly	• longer overall life due to recharging • cheaper in the long run	• very short shelf-life • needs to be recharged often

Section 1: Activity 3

1. Oxidation occurs at the anode.

2. Reduction occurs at the cathode.

3. Cations are positive ions that migrate to the cathode. Anions are negative ions that migrate to the anode.

4. a. Two examples of a porous boundary are a salt bridge and a porous cup.

   b. The boundary must be porous to allow charge transfer in the form of ions.

   c. The porous boundary is represented by double vertical lines, ‖, in the cell notation.

5. Chemists separate voltaic cells into two half-cells to make it easier to study them and to keep the ions from randomly mixing before the desired cell operation.

6. a. vi b. iv c. i d. ii e. iii f. v

7. Your results should be similar to the following:

Type of Boundary	Cathode	Anode	Initial Potential Difference	Observations (after 5 min)	Observations (after 24 h)
None	$Cu_{(s)}$	$Zn_{(s)}$	0.81 V	The potential has dropped to zero.	No observable change is seen.
Salt bridge	$Cu_{(s)}$	$Zn_{(s)}$	0.89 V	0.89 V – Some black colour is on zinc.	Material collects on the cathode. The blue solution disappears. The zinc looks dark.
Porous cup	$Cu_{(s)}$	$Zn_{(s)}$	0.89 V	0.89 V – Some black colour is on zinc.	Material collects on the copper. The blue solution disappears. The zinc looks dark.

8. The values obtained for the salt bridge and porous cup were the same. However, the $NaNO_{3(aq)}$ electrolyte gave a lower reading. This is because the electrolyte is different than in the other two cells and should therefore be expected to yield a different potential (as seen in Investigation 13.2).

9. The electrolyte was made of $NaNO_{3(aq)}$ so an unwanted redox side reaction wouldn't occur in the solution. The sodium and nitrate ions in $NaNO_{3(aq)}$ would only be spectator ions.

10. a. If reduction occurs at the cathode, then oxidation occurs at the **anode**.

    b. Reduction means gain of electrons, so electrons must flow to the **cathode** from the **anode**. Electrons flow through the wire because that is how charge flows through metals. Charge can only flow through solution by using **ions** to transfer the charge.

    c. Cations are usually reduced, so they must flow to the **cathode** to gain electrons. When cations are reduced, a metal usually forms which means the mass at the cathode usually **increases**. Likewise, the anode usually **loses** mass because as oxidation occurs the metal electrode forms cations.

    d. When cations are reduced at the cathode, there become more anions than cations in the reduction half-cell. Since like charges repel, the anions will leave the **cathode (reduction)** half-cell and migrate to the **anode (oxidation)** half-cell. Likewise, the cations at the **anode** must be in excess which means they will leave the **anode (oxidation)** half-cell.

11. reduction (gain $1\,e^-$)

$$\overset{0}{Zn_{(s)}} + 2\,\overset{+4\ -2}{MnO_{2(s)}} + 2\,\overset{-3\ +1\ -1}{NH_4Cl_{(aq)}} \rightarrow \overset{+2\ -1}{ZnCl_{2(aq)}} + \overset{+3\ -2}{Mn_2O_{3(s)}} + 2\,\overset{-3\ +1}{NH_{3(aq)}} + \overset{+1\ -2}{H_2O_{(l)}}$$

oxidation (lose $2\,e^-$)

Since Mn is being reduced, the carbon electrode surrounded by $MnO_{2(s)}$ is the cathode. Since Zn is being oxidized, the $Zn_{(s)}$ casing is the anode. The spacer is the porous membrane.

12. **Textbook question 16:**

    a. cathode
    b. anode
    c. anode
    d. cathode

    Textbook question 17:

    a. Ions move through the porous barrier to maintain electrical neutrality in the cell.

    b. Cations move toward the cathode because they are attracted by the excess negative charge left in solution around the cathode.

    Anions move toward the anode because they are attracted by the excess positive charge being produced in solution around the anode.

Textbook question 18:

An inert electrode is a solid conducting substance that won't interfere with the desired reaction. It is used when the strongest oxidizing agent or reducing agent is not a solid and cannot act as the cathode or anode.

Textbook question 19:

The aqueous solution in a salt bridge contains an inert electrolyte so that it will conduct a current without interfering in the reaction. One example could be $Na_2SO_{4(aq)}$.

Textbook question 20:

a. SOA OA

$Ag_{(s)} \mid Ag^+_{(aq)} \parallel Zn^{2+}_{(aq)} \mid Zn_{(s)}$

RA SRA

(**Note:** The presence of $H_2O_{(l)}$ is implied in each case by the aqueous states.)

Cathode Half-Reaction: SOA is the reactant.	**Anode Half-Reaction:** SRA is the reactant.
$Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$	$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2\ e^-$
Net Reaction: $2\ Ag^+_{(aq)} + Zn_{(s)} \rightarrow 2\ Ag_{(s)} + Zn^{2+}_{(aq)}$	

b. SOA OA

$Cu_{(s)} \mid Cu^{2+}_{(aq)} \parallel Zn^{2+}_{(aq)} \mid Zn_{(s)}$

RA SRA

Cathode Half-Reaction: SOA is the reactant.	**Anode Half-Reaction:** SRA is the reactant.
$Cu^{2+}_{(aq)} + 2\ e^- \rightarrow Cu_{(s)}$	$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2\ e^-$
Net Reaction: $Cu^{2+}_{(aq)} + Zn_{(s)} \rightarrow Cu_{(s)} + Zn^{2+}_{(aq)}$	

c. OA SOA OA

$Sn_{(s)} \mid Sn^{2+}_{(aq)} \parallel Cr_2O_7^{2-}{}_{(aq)}, H^+_{(aq)} \mid C_{(s)}$

SRA RA

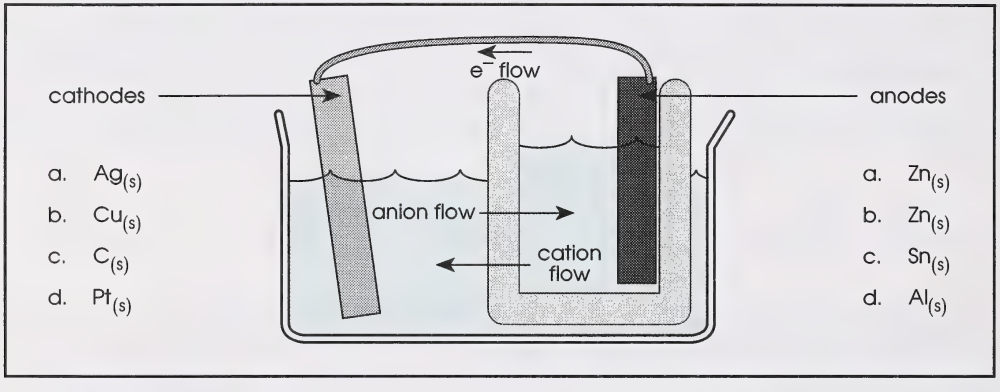
Cathode Half-Reaction: SOA is the reactant.	**Anode Half-Reaction:** SRA is the reactant.
$Cr_2O_7^{2-}{}_{(aq)} + 14\ H^+_{(aq)} + 6\ e^- \rightarrow 2\ Cr^{3+}_{(aq)} + 7\ H_2O_{(l)}$	$Sn_{(s)} \rightarrow Sn^{2+}_{(aq)} + 2\ e^-$
Net Reaction: $Cr_2O_7^{2-}{}_{(aq)} + 14\ H^+_{(aq)} + 3\ Sn_{(s)} \rightarrow 2\ Cr^{3+}_{(aq)} + 7\ H_2O_{(l)} + 3\ Sn^{2+}_{(aq)}$	

d. $\underset{\text{SRA}}{Al_{(s)}} \mid Al^{3+}_{(aq)} \parallel Na^{+}_{(aq)}, Cl^{-}_{(aq)}, \underset{\text{RA}}{O_{2(g)}, H_2O_{(l)}} \mid Pt_{(s)}$ (OA; SOA, OA)

Cathode Half-Reaction: SOA is the reactant.	**Anode Half-Reaction:** SRA is the reactant.
$O_{2(g)} + 2\,H_2O_{(l)} + 4\,e^- \rightarrow 4\,OH^-_{(aq)}$	$Al_{(s)} \rightarrow Al^{3+}_{(aq)} + 3\,e^-$
Net Reaction: $3\,O_{2(g)} + 6\,H_2O_{(l)} + 4\,Al_{(s)} \rightarrow 4\,Al(OH)_{3(s)}$	

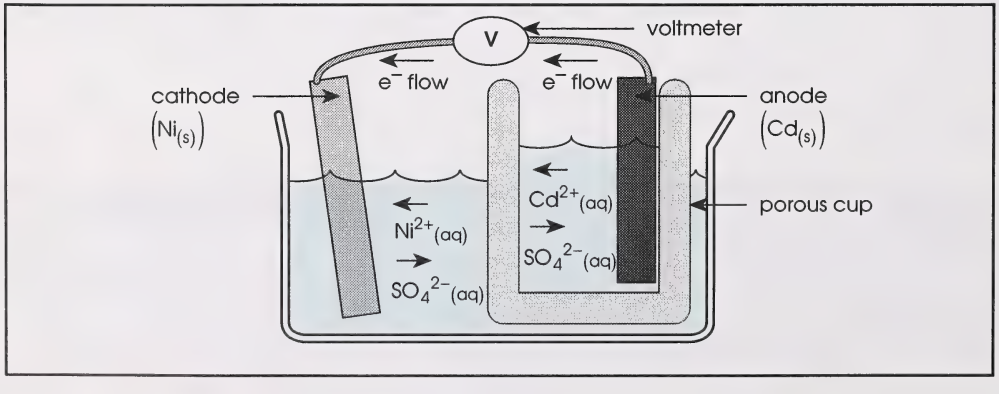
Note: The $Al^{3+}_{(aq)}$ and $OH^-_{(aq)}$ ions would combine to produce the precipitate $Al(OH)_{3(s)}$.

The following cell diagram applies for all cells a-d. Assume that the appropriate electrolytes are in their correct half-cell solutions in each case. (Oxygen gas would be bubbled around the platinum cathode in cell d.)



Textbook question 21:

Your diagram of a voltaic cell could be similar to the following:



Section 1: Activity 4

1. The strongest side usually wins a tug-of-war.

2. a. ii b. i c. iii d. ii e. iv f. i

3. Since it is not possible to determine when an electron has zero potential energy, one was arbitrarily chosen. Measurements can then be made with reference to the zero point.

4. Choosing an arbitrary zero point for a standard reference half-cell does not affect the voltage of a cell because the cell potential depends only on how far apart the potential energies of the electrons are. The relative difference between half-cell $E°$ values is the important value, not the actual $E°$ values themselves.

5. A positive $E°$ reduction potential value means that the oxidizing agent attracts electrons stronger than hydrogen ions (hydrogen half-cell). A negative $E°$ value means that the oxidizing agent attracts electrons weaker than hydrogen ions.

6. **Textbook question 27:**

   To answer this question, you first have to ask yourself, how does the $E°$ value for the standard lithium half-cell, –3.04 V, have to change in order to become 0.00 V? To change –3.04 V to 0.00 V, you have to **add** 3.04 V. Therefore, the new $E°$ values for the copper and zinc half-cells would be found by adding 3.04 V to each of their values.

   copper: $+0.34\text{ V} + 3.04\text{ V} = +3.38\text{ V}$
   zinc: $-0.76\text{ V} + 3.04\text{ V} = +2.28\text{ V}$

   The following table shows the new $E°_r$ (lithium) for copper and zinc. Notice that the relative positions of the $E°$ values do not change.

Reduction Half-Reactions	REFERENCE HALF-CELL		Relative Differences
	Hydrogen	Lithium	
$Cu^{2+}_{(aq)} + 2\,e^- \rightarrow Cu_{(s)}$	$E° = +0.34$ V	$E° = +3.38$ V	+1.10 V; +3.38 V
$Zn^{2+}_{(aq)} + 2\,e^- \rightarrow Zn_{(s)}$	$E° = -0.76$ V	$E° = +2.28$ V	+2.28 V
$Li^{+}_{(aq)} + e^- \rightarrow Li_{(s)}$	$E° = -3.04$ V	$E° = 0.00$ V	

7. **Textbook question 23:**

a.

Cathode: (Reactant is SOA)	$Sn^{2+}_{(aq)} + 2\,e^- \rightarrow Sn_{(s)}$
Anode: (Reactant is SRA)	$Cr_{(s)} \rightarrow Cr^{2+}_{(aq)} + 2\,e^-$
Net Reaction:	$Sn^{2+}_{(aq)} + Cr_{(s)} \rightarrow Sn_{(s)} + Cr^{2+}_{(aq)}$

$\Delta E° = E°_r(\text{cathode}) - E°_r(\text{anode})$
$= -0.14\text{ V} - (-0.91\text{ V})$
$= +0.77\text{ V}$

b.

Cathode: (Reactant is SOA)	$SO_4^{2-}_{(aq)} + 4\,H^+_{(aq)} + 2\,e^- \rightarrow H_2SO_{3(aq)} + H_2O_{(l)}$
Anode: (Reactant is SRA)	$Co_{(s)} \rightarrow Co^{2+}_{(aq)} + 2\,e^-$
Net Reaction:	$SO_4^{2-}_{(aq)} + 4\,H^+_{(aq)} + Co_{(s)} \rightarrow H_2SO_{3(aq)} + H_2O_{(l)} + Co^{2+}_{(aq)}$

$\Delta E° = E°_r(\text{cathode}) - E°_r(\text{anode})$
$= +0.17\text{ V} - (-0.28\text{ V})$
$= +0.45\text{ V}$

c.

Cathode: (Reactant is SOA)	$O_{2(g)} + 2\,H_2O_{(l)} + 4\,e^- \rightarrow 4\,OH^-_{(aq)}$
Anode: (Reactant is SRA)	$H_{2(g)} + 2\,OH^-_{(aq)} \rightarrow 2\,H_2O_{(l)} + 2\,e^-$
Net Reaction:	$O_{2(g)} + \cancel{2\,H_2O_{(l)}} + 2\,H_{2(g)} + \cancel{4\,OH^-_{(aq)}} \rightarrow \cancel{4\,OH^-_{(aq)}} + 2\,\cancel{4}\,H_2O_{(l)}$
Cancel and **Simplify to:**	$O_{2(g)} + 2\,H_{2(g)} \rightarrow 2\,H_2O_{(l)}$

$\Delta E° = E°_r(\text{cathode}) - E°_r(\text{anode})$
$= +0.40\text{ V} - (-0.83\text{ V})$
$= +1.23\text{ V}$

Textbook question 24:

Except where indicated, the standard cell notations are written in the form:

anode (**SRA**) | anode electrolyte || cathode electrolyte (**SOA**) | cathode

a. $Pb_{(s)} \left| Pb^{2+}_{(aq)} \right\| Cu^{2+}_{(aq)} \left| Cu_{(s)} \right.$

$$\begin{aligned}\Delta E^\circ &= E^\circ_r(\text{cathode}) - E^\circ_r(\text{anode}) \\ &= +0.34\text{ V} - (-0.13\text{ V}) \\ &= +0.47\text{ V}\end{aligned}$$

b. $Zn_{(s)} \left| Zn^{2+}_{(aq)} \right\| Ni^{2+}_{(aq)} \left| Ni_{(s)} \right.$

$$\begin{aligned}\Delta E^\circ &= E^\circ_r(\text{cathode}) - E^\circ_r(\text{anode}) \\ &= -0.26\text{ V} - (-0.76\text{ V}) \\ &= +0.50\text{ V}\end{aligned}$$

c. $Pt_{(s)}\ (\text{or } C_{(s)}) \left| H^+_{(aq)},\ H_{2(g)}\ \textbf{(SRA)} \right\| \textbf{(SOA)}\ Fe^{3+}_{(aq)},\ Fe^{2+}_{(aq)} \left| Pt_{(s)}\ (\text{or } C_{(s)}) \right.$

$Fe^{2+}_{(aq)}$ was needed because both reactants and products in the reaction must be included. The cathode had to be inert so the following reaction would not occur: $Fe_{(s)} + 2\ Fe^{3+}_{(aq)} \rightarrow 3\ Fe^{2+}_{(aq)}$

$$\begin{aligned}\Delta E^\circ &= E^\circ_r(\text{cathode}) - E^\circ_r(\text{anode}) \\ &= +0.77\text{ V} - (0.00\text{ V}) \\ &= +0.77\text{ V}\end{aligned}$$

Textbook question 25:

Since $E^\circ_r\left(Au_{(s)}\right) = +1.50$ V and the gold cell is the cathode, the $In_{(s)}$ must be 1.84 V lower than gold's electrode potential. Therefore, $E^\circ_r\left(In_{(s)}\right)$ must be -0.34 V. This can be proven from the equation

$$\Delta E^\circ = E^\circ_r(\text{cathode}) - E^\circ_r(\text{anode})$$

(rearranging yields) $E^\circ_r(\text{anode}) = E^\circ_r(\text{cathode}) - \Delta E^\circ$

$$\begin{aligned}&= +1.50\text{ V} - (+1.84\text{ V}) \\ E^\circ_r(\text{In}) &= -0.34\text{ V}\end{aligned}$$

Textbook question 26:

From the reaction, $Ag^+_{(aq)}$ is reduced, so $X_{(s)}$ must be oxidized. Therefore, $X_{(s)}$ is the anode and must be 1.08 V below the $Ag^+_{(aq)}$ reduction half-reaction on the redox table. Knowing $E^\circ_r\left(Ag^+_{(aq)}\right) = 0.80$ V, $E^\circ_r\left(X_{(s)}\right) = 0.80\text{ V} - 1.08\text{ V}$. Thus, $E^\circ_r\left(X_{(s)}\right) = -0.28$ V. From the table of reduction half-reactions, the substance with an $E^\circ_r = -0.28$ V is $Co_{(s)}$. Therefore $X_{(s)}$ is $Co_{(s)}$, and $X^{2+}_{(aq)}$ is $Co^{2+}_{(aq)}$.

Substituting into the equation $\Delta E^\circ = E^\circ_r(\text{cathode}) - E^\circ_r(\text{anode})$ would yield the same result.

8. You do not multiply the $E°$ values because the electrode potential is the potential energy per unit charge where 1 V = 1 J/C. If you doubled the number of electrons, you would also double the total potential energy. The ratio of potential energy to charge would therefore remain unchanged.

9. The following explanations expand on the answers given for questions 17, 19, and 20 on page 517 of *Nelson Chemistry*.

Textbook question 17:

The $E°_r$ of +0.40 V is the largest, so the $O_{2(g)}$ equation represents the reduction half-reaction (the cathode). That means the other half-reaction equation must be reversed. Therefore, the net reaction is

Cathode: $2\left(O_{2(g)} + 2\,H_2O_{(g)} + 4\,e^- \rightarrow 4\,OH^-_{(l)}\right)$

Anode: $10\,OH^-_{(l)} + CH_{4(g)} \rightarrow CO_3{}^{2-}_{(l)} + 7\,H_2O_{(g)} + 8\,e^-$

Net Reaction (after cancelling): $2\,O_{2(g)} + 2\,OH^-_{(l)} + CH_{4(g)} \rightarrow CO_3{}^{2-}_{(l)} + 3\,H_2O_{(g)}$

Substituting the $E°_r$ values into the net reaction equation,

$$\begin{aligned}\Delta E° &= E°_r(\text{cathode}) - E°_r(\text{anode}) \\ &= +0.40\text{ V} - (+0.17\text{ V}) \\ &= +0.23\text{ V}\end{aligned}$$

Textbook question 19:

Since $Cu^+_{(aq)}$ is being reduced, its reduction would occur at the cathode. Therefore, the reduction potential should be 1.28 V above −0.76 V. In other words, $E°_r\left(Cu^+_{(aq)}\right) = +0.52\text{ V}$.

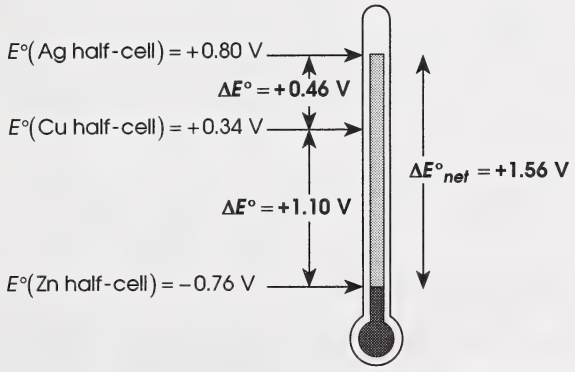
Substituting into the equation $\Delta E° = E°_r(\text{cathode}) - E°_r(\text{anode})$ would yield the same result.

Textbook question 20:

Since $Cu^{2+}_{(aq)}$ is being reduced, $X_{(s)}$ is being oxidized. Therefore the $E°$ value for $X_{(s)}$ it is 0.48 V below the $Cu^{2+}_{(aq)}$ reduction half-reaction equation whose $E°_r = +0.34\text{ V}$. In other words, $E°_r\left(X^+_{(aq)}\right) = -0.14\text{ V}$.

Substituting into the equation, $\Delta E° = E°_r(\text{cathode}) - E°_r(\text{anode})$ would yield the same result. The table of redox half-reaction equations reveals that the pair $X^{2+}_{(aq)} \,|\, X_{(s)}$ is $Sn^{2+}_{(aq)} \,|\, Sn_{(s)}$.

10. You could calculate the cell potential of the $Cu_{(s)} \mid Cu^{2+}_{(aq)} \parallel Ag^{+}_{(aq)} \mid Ag_{(s)}$ cell and the cell potential of the $Zn_{(s)} \mid Zn^{2+}_{(aq)} \parallel Cu^{2+}_{(aq)} \mid Cu_{(s)}$ cell, and then add the two cell potentials to get the voltmeter reading. You add the values because the two half-cells are in series. However, if you compare Figure 13.22 on page 406 of *Nelson Chemistry* with Figure 13.16 on page 399, you might notice that the $Cu_{(s)}$ and $Cu^{2+}_{(aq)}$ solutions of Figure 13.22 look remarkably like the salt bridge used in Figure 13.16. In fact they behave exactly the same – both are used just to transfer charge. Knowing this, you could simply calculate the cell potential of a $Zn_{(s)} \mid Zn^{2+}_{(aq)} \parallel Ag^{+}_{(aq)} \mid Ag_{(s)}$ cell and get your answer more quickly. This is illustrated in the following diagram.



From the diagram it can be seen that using the equation $\Delta E^\circ = E^\circ_r(\text{cathode}) - E^\circ_r(\text{anode})$ would also yield $\Delta E^\circ = +1.56$ V.

Section 1: Activity 5

1. The answer is **d.** and **h.** Rust remover only removes corrosion once it has occurred but will not protect against further corrosion. Adding salt forms an electrolyte in solution that will increase corrosion. Most rust protection methods work by creating a barrier that prevents water and oxygen from coming in contact with the iron. This is true in the cases of paint, a plastic coating, an oxide coating, and metal plating.

2. Brass is mostly copper and has a higher E° than iron, so when moisture comes in contact where the fastener touches the screen, a simple cell is formed with iron as the anode, brass as the cathode, and the ions in water as the electrolyte. Therefore, while brass won't corrode, the iron will rust around the fastener and the screen will soon need to be replaced.

   Zinc and aluminum are both stronger reducing agents than iron, so they would be the anodes of the simple cell that forms, while iron is the cathode. Therefore, if any corrosion occurs, it will be the fastener only that will need replacing.

   If iron is used, the water will get in the cracks where the fastener meets the screen and both will rust, as explained in Figure 13.15 on page 398 of *Nelson Chemistry*.

   Therefore, the best fasteners would be aluminum or zinc; the poorest fasteners would be brass and iron.

3. a. False. The half-reaction is reduction, but the potential given is an oxidation potential. The potential should be -0.36 V.

   b. True

   c. True

   d. False. The half-reaction is reduction, but the potential given is an oxidation potential. The potential should be $+1.07$ V.

4. If M is more easily oxidized than Q, then M has a higher oxidation potential than Q, so the order from highest to lowest is M, Q.

   If P will cathodically protect Q, then P is more easily oxidized, which means the order of this pair is P, Q.

   If T is the sacrificial anode for P, then T is more easily oxidized than P, which means the order is T, P.

   And if T is the cathode, then M is the anode and is oxidized, which means the order for this pair is M, T.

   By logic, now, if P is higher than Q (second clue), and T is higher than P (third clue), then T is higher than Q and the order for these three is T, P, Q. But M is higher than T so the overall order from strongest reducing agent (highest oxidation potential) to lowest is M, T, P, Q.

   - M above Q
   - P above Q
   - T above P
   - M above T

   → M ← SRA – highest oxidation potential
   T
   P
   Q

5. The cathode is $Fe_{(s)}$ because it has a higher reduction potential than the anode $Mg_{(s)}$. Electrons will flow to the $Fe_{(s)}$ from $Mg_{(s)}$ along the copper wire. The electrolyte is the many different ions found in the moisture of the soil.

6. Your predictions should be similar to the following (all cells here are written cathode first and anode last):

   Cell 1: $Cu_{(s)} \left| Cu^{2+}_{(aq)} \right\| Pb^{2+}_{(aq)} \left| Pb_{(s)} \right.$

   $\Delta E^\circ = +0.34 \text{ V} - (-0.13 \text{ V})$

   $= +0.47 \text{ V}$

   Cell 2: $Cu_{(s)} \left| Cu^{2+}_{(aq)} \right\| Zn^{2+}_{(aq)} \left| Zn_{(s)} \right.$

   $\Delta E^\circ = +0.34 \text{ V} - (-0.76 \text{ V})$

   $= +1.10 \text{ V}$

   Cell 3: $Pb_{(s)} \left| Pb^{2+}_{(aq)} \right\| Zn^{2+}_{(aq)} \left| Zn_{(s)} \right.$

   $\Delta E^\circ = -0.13 \text{ V} - (-0.76 \text{ V})$

   $= +0.63 \text{ V}$

7. Your procedure should be similar to the following.

- Polish the $Cu_{(s)}$, $Pb_{(s)}$, and $Zn_{(s)}$ with steel wool.
- Fill one 50 mL or 100 mL beaker with 25 mL to 35 mL of 0.10 mol/L $CuSO_{4(aq)}$, another beaker with 0.10 mol/L $Pb(NO_3)_{(aq)}$, and a third with 0.10 mol/L $Zn(NO_3)_{2(aq)}$.
- Fill a U-tube with a 0.10 mol/L solution of $NaNO_{3(aq)}$ and plug both ends with cotton batting.
- Place each metal in its corresponding ionic solution.
- Set up a cell between the copper half-cell and the lead half-cell as shown in Figure 13.13 (b) on page 396 of *Nelson Chemistry*. Connect the red lead to the cathode and the black lead to the anode. (The salt bridge is the U-tube containing $NaNO_{3(aq)}$.)
- Take a voltage reading.
- Repeat the previous two steps for the other two cells, using a fresh salt bridge each time.

8. Your observations should be similar to the following:

Cell (Anode/Cathode)*	Potential Difference (V ± 0.01)
$Pb_{(s)} \mid Pb^{2+}_{(aq)} \parallel Cu^{2+}_{(aq)} \mid Cu_{(s)}$	+0.46 V
$Zn_{(s)} \mid Zn^{2+}_{(aq)} \parallel Cu^{2+}_{(aq)} \mid Cu_{(s)}$	+0.96 V
$Zn_{(s)} \mid Zn^{2+}_{(aq)} \parallel Pb^{2+}_{(aq)} \mid Pb_{(s)}$	+0.52 V

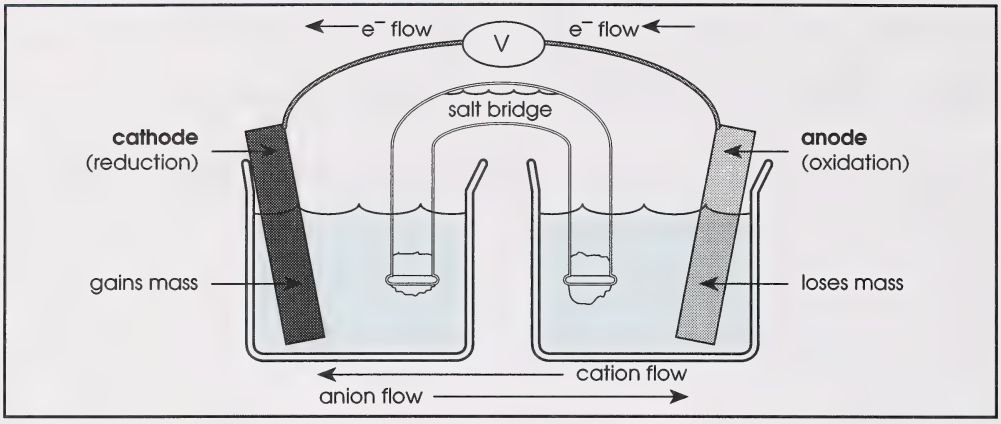
* The anode is written first and the cathode is written last.

9. The results should be slightly lower than predicted because the predicted cell potentials are for 1.0 mol/L solutions and SATP conditions. The cells made in this experiment are made from only 0.10 mol/L solutions and are not run under SATP conditions. If the concentration of the solution drops, it is only reasonable to expect a drop in cell potential because it is more difficult to transfer the charge through the solution.

Section 1: Follow-up Activities

Extra Help

1.



2. Voltaic cells consist of three main parts – the **cathode**, **anode**, and **electrolyte**. In the cell, electrons move from the **anode** to the **cathode**. The solution between the two electrodes is an **electrolyte**, because it conducts electricity and is needed to complete the **circuit**. As the electrons move through the wire, they lose **energy**. The amount of energy they lose depends on how far apart the **electrode potentials** are. The further apart they are, the greater the **net cell potential** of the cell is. If the **strongest oxidizing agent** is higher than the **strongest reducing agent** on a table of reduction half-reactions, then the reaction is **spontaneous** and the net cell potential is **positive**. An example of this type of cell can be used to start a robot.

3. a. ii
   c. iii
   b. iv
   d. i

4. Since magnesium is a stronger reducing agent than iron, it will be more easily oxidized. Since oxidation occurs at the anode (An Ox), magnesium will become the sacrificial anode and will cathodically protect the eavestrough for a small area near where the magnesium is attached. The magnesium will corrode as a result.

Enrichment

1. Your answer should be similar to the following.

   Batteries are important to society because they provide portable energy to the consumer. This portable energy may be used for entertainment, as with a portable CD player, or in business, as with a calculator, laptop computer, and various field testing kits. They may also be used in construction and general appliances, such as with a portable drill, screwdriver, or even a hand-held vacuum cleaner. But one of the biggest challenges for the battery is in transportation – through the development of the electric car.

   While batteries will probably be used for a very long time to come, one opinion is that disposable batteries will become more and more obsolete as rechargeable batteries become better and better, even though disposables are cheap to make. However, even rechargeable batteries may become less desirable as the technology of fuel cells improves.

But batteries will not solve all of society's energy needs. The rate at which they produce energy is often too low as is their energy density. In addition, even though battery lifetimes may be extended, they eventually need replacing which is expensive. Therefore, there will still be the problem of disposal. Currently in Canada, the only serious effort in the recycling of batteries is in reclaiming the lead from regular lead-acid car batteries.

In twenty to fifty years fuel cells may become very popular as an alternative energy source for many larger users of electricity such as small generators, electric cars, and even laptop computers. The concern for the environment will only encourage this trend, as the burning of fossil fuels is discouraged more and more.

The following table summarizes some information on the various types of batteries available, and their uses:

Battery Type	Pros	Cons	Examples of Uses
dry cell	cheapest	short life; cold hampers it	general purpose (flashlights, toys, etc.)
alkalines	longer lasting than dry cell	more expensive	long-drain general purpose (flashes, etc. – good in cold)
nickel-cadmium	high current available; lower voltage	has memory effect; fast discharge	camcorders, power tools, portable phones, and other small items
nickel-metal hydride	longer life than Ni-Cads	give only low current	camcorders, laptop computers, etc.
lead-acid	cheapest rechargeable	heavy in weight	automobiles, electric cars, laptop computers
sodium-sulphur	versatile; rechargeable; medium range life	high operating temperature	electric cars
lithium-polymer	high energy density	technology not fully developed	small appliances, electric cars
phosphoric acid fuel cell	versatile	below average battery efficiency	small generators, electric cars, small appliances
zinc-air fuel cell	very long life per charge	low number of recharges	computers, field test devices
alkaline-potassium hydroxide fuel cell	70% efficient	extremely expensive; high operating temperature	space shuttle, military

The preceding table gives a sample only of the various types of batteries and fuel cells available and where they may be used. Most of these have been developed recently, and many more are being developed. This suggests that industry anticipates a high demand for batteries and fuel cells in the future.

2. While recycling is a trend that seems to be growing, battery recycling is not yet catching on. The main reason is the difficulty of recycling batteries. Since a battery never fully loses its charge, this needs to be taken into account in the recycling or disposing process. In addition, the batteries are constructed in such a way that it is hard to separate the materials.

   As a result, the main type of battery that people are interested in recycling is the regular lead-acid car battery, and the main substance they are interested in is the lead. The plastic is also kept for recycling, but the sulphuric acid in the battery must be neutralized. In addition, the solution contains a high amount of lead ions – a heavy metal. For this reason many special precautions must be taken in the disposal of the solutions, and the cost of cleaning up the site is expensive. For example, in Calgary, a battery recycling firm that recycled batteries from 1979 until 1982 left soil heavily contaminated with lead ions. In 1993, a study was started to find ways to clean up the site. In the meantime, the area has been fenced off and the land periodically sprayed to keep the contamination from spreading.

   But what are the alternatives to battery recycling? Some companies have looked at ways of extending the life of batteries or finding ways to recharge primary cells, but these efforts are limited. A more feasible alternative may be to cut back on battery consumption by the consumer, using alternative sources for convenient electrochemical energy.

3. To determine the difference between a dry cell, alkaline cell, and a rechargeable cell, the easiest way is to find a device, such as a flashlight, that will operate on all three types of cells. Use the device to test each battery by turning the item on to see how long it lasts. To determine the rate at which the battery loses its power output, you could measure the speed at which the device is operating. For example, with a flashlight you could use a light meter to measure the light intensity at regular intervals. You should repeat the experiment a number of times to get consistent results.

   Your results should show that a dry cell discharges quickly, but at a steady rate. An alkaline cell lasts the longest and delivers a steady current until near the end of its life, and a rechargeable cell gives a steady current until it dies very suddenly. The rechargeable's life-span varies because it may not have been fully discharged or charged. This creates a memory effect (the battery is no longer able to discharge fully) and this causes the battery performance to go down.

Section 2: Activity 1

1. a. The strongest oxidizing agent in the reaction is $Pb^{2+}_{(aq)}$, while the strongest reducing agent is $Fe_{(s)}$.

      Since $Fe^{2+}_{(aq)}$ is not one of the reactants and an electrolyte is still present in the solution, the cell will still operate.

   b. If the iron is replaced with copper, then the strongest oxidizing agent is lower on the table of redox half-reactions than the strongest reducing agent. Therefore the predicted reaction is non-spontaneous which means the cell will not operate.

   c. If the solutions are allowed to mix, the $Pb^{2+}_{(aq)}$ can come into direct contact with the iron anode. Therefore electron transfer will occur directly rather than through the wire, and the cell will not work.

2. Since the reaction becomes non-spontaneous if the iron anode is replaced with copper, electron transfer will occur only if forced to do so. Therefore, there is no longer any need to separate the half-reactions.

3. The answer is **C**. One difference between electrolytic and voltaic cells is that the electrolytic cell must have a non-spontaneous reaction.

4. **COMPARING THE VOLTAIC CELL WITH THE ELECTROLYTIC CELL**

Similarities	Differences	
	Voltaic	Electrolytic
• Reduction occurs at the cathode. • Oxidation occurs at the anode. • Cations move to the cathode. • Anions move to the anode. • Electrons go to the cathode. • The cathode often gains mass. • The anode often loses mass.	• spontaneous • chemical → electrical energy energy • $\Delta E°$ positive • no power supply • porous membrane required • SOA higher than SRA on redox table	• non-spontaneous • electrical → chemical energy energy • $\Delta E°$ negative • needs power supply • no porous membrane required • SRA higher than SOA on redox table

5. a. Besides having a non-spontaneous reaction occur in an electrolytic cell, the main difference between a voltaic cell and an electrolytic cell is that the electrolytic cell requires an **external power source**.

   b. One way to recognize whether a cell is voltaic or electrolytic is by looking up the reactants in your table of redox half-reactions. If the strongest oxidizing agent is at the same level or lower than the strongest reducing agent, then the cell is an **electrolytic** cell. Another way is to see if the cell has an external **power source**.

6. **Textbook question 13:**

   See textbook answer 13 on page 517 of *Nelson Chemistry*.

   Textbook question 21:

   See textbook answer 21 on page 517 of *Nelson Chemistry*. The following is an explanation of the answer.

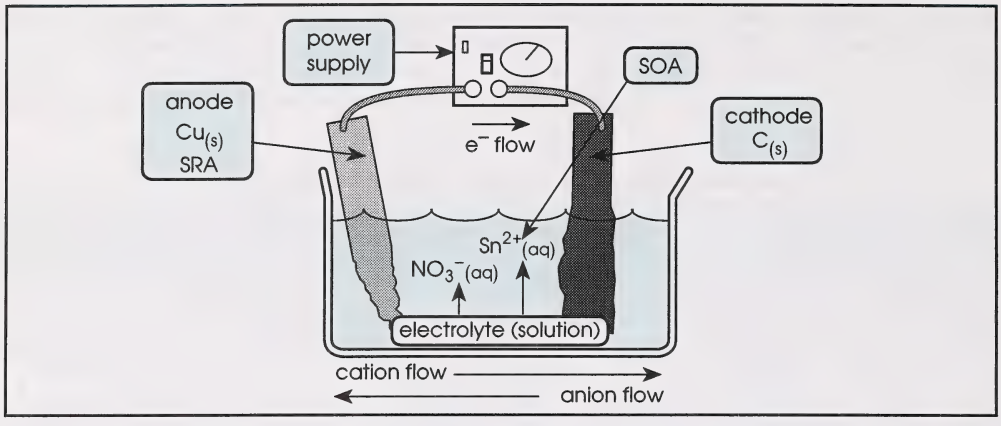
   a. The SOA is $Cd^{2+}_{(aq)}$ which is lower than the SRA of $H_2O_{(l)}$. Therefore the reaction is non-spontaneous and an external voltage is needed.

   b. The SOA is $Fe^{3+}_{(aq)}$ which is higher than the SRA of $I^-_{(aq)}$. Therefore the reaction is spontaneous and an external voltage is not needed.

   c. The SOA is $Fe^{3+}_{(aq)}$ which is higher than the SRA of $Sn^{2+}_{(aq)}$. Therefore the reaction is spontaneous and an external voltage is not needed.

   d. The SOA is $Zn^{2+}_{(aq)}$ which is lower than the SRA of $I^-_{(aq)}$. Therefore the reaction is non-spontaneous and an external voltage is needed.

7.



8. Your evidence table and observations should be similar to the following:

Observations	Before Connection	After Connection
Cathode	• The red litmus stayed red. • The blue litmus stayed blue. • The solution was clear and colourless.	• The red litmus turned blue. • The blue litmus stayed blue. • The solution was clear and colourless. • Bubbles were produced.
Anode	• The red litmus stayed red. • The blue litmus stayed blue. • The solution was clear and colourless.	• The red litmus stayed red. • The blue litmus stayed blue. • The solution turned brown around the electrode. • The anode collected some dark solid.

9. Since the only species present were $K^+_{(aq)}$, $I^-_{(aq)}$, and $H_2O_{(l)}$, the strongest oxidizing agent which reacts at the cathode was $H_2O_{(l)}$ and its half-reaction is $2\,H_2O_{(l)} + 2\,e^- \rightarrow H_{2(g)} + 2\,OH^-_{(aq)}$. Likewise, the strongest reducing agent which reacts at the anode is $I^-_{(aq)}$ and its half-reaction is $2\,I^-_{(aq)} \rightarrow I_{2(s)} + 2\,e^-$.

From the litmus results, $OH^-_{(aq)}$ was produced at the cathode but no $H^+_{(aq)}$ or $OH^-_{(aq)}$ was produced at the anode. From the predicted half-reaction and litmus paper observations, the gas produced at the cathode was most likely $H_{2(g)}$. The red-brown substance produced at the anode was most likely dissolved iodine. The dark solid was most likely solid iodine. If an anode solution sample was tested for iodine (halogen diagnostic test), the purple solvent layer (trichlorotrifluoroethane) would confirm the presence of iodine.

10. **Cathode** (reduction): $2\,H_2O_{(l)} + 2\,e^- \rightarrow H_{2(g)} + 2\,OH^-_{(aq)}$

Anode (oxidation): $2\,I^-_{(aq)} \rightarrow I_{2(s)} + 2\,e^-$

Net Reaction: $2\,H_2O_{(l)} + 2\,I^-_{(aq)} \rightarrow H_{2(g)} + 2\,OH^-_{(aq)} + I_{2(s)}$

11. **Textbook question 28:**

As before with voltaic cells, the following two formulas can be used to calculate the net cell potentials:

$$\Delta E^\circ = E^\circ_r(\text{cathode}) - E^\circ_r(\text{anode}) \quad \textbf{or} \quad \Delta E^\circ = E^\circ_r(\text{cathode}) + E^\circ_{ox}(\text{anode})$$

a. **SOA** **OA**
$Ni^{2+}_{(aq)}$, $I^-_{(aq)}$, $H_2O_{(l)}$
SRA **RA**

Cathode: $Ni^{2+}_{(aq)} + 2\,e^- \rightarrow Ni_{(s)}$ $\quad E^\circ_r = -0.26$ V

Anode: $2\,I^-_{(aq)} \rightarrow I_{2(s)} + 2\,e^-$ $\quad E^\circ_{ox} = -0.54$ V

Net Reaction: $Ni^{2+}_{(aq)} + 2\,I^-_{(aq)} \rightarrow Ni_{(s)} + I_{2(s)}$ $\quad \Delta E^\circ = -0.80$ V

or, $\Delta E^\circ = -0.26\text{ V} - (+0.54\text{ V})$
$= -0.80$ V The potential difference is -0.80 V.

OA **SOA**
b. $Na^+_{(aq)}$, $OH^-_{(aq)}$, $H_2O_{(l)}$ ← **Note:** Water now participates as the SOA.
SRA **RA**

Cathode: $2\left(2\,H_2O_{(l)} + 2\,e^- \rightarrow H_{2(g)} + 2\,OH^-_{(aq)}\right)$ $\quad E^\circ_r = -0.83$ V

Anode: $4\,OH^-_{(aq)} \rightarrow O_{2(g)} + 2\,H_2O_{(l)} + 4\,e^-$ $\quad E^\circ_{ox} = -0.40$ V

Total Reaction: $2\,\cancel{4\,H_2O_{(l)}} + \cancel{4\,OH^-_{(aq)}} \rightarrow 2\,H_{2(g)} + \cancel{4\,OH^-_{(aq)}} + O_{2(g)} + \cancel{2\,H_2O_{(l)}}$

After cancelling, the net reaction is $2\,H_2O_{(l)} \rightarrow 2\,H_{2(g)} + O_{2(g)}$ $\quad \Delta E^\circ = -1.23$ V

or, $\Delta E^\circ = -0.83\text{ V} - (+0.40\text{ V})$
$= -1.23$ V

Textbook question 29:

a. SOA OA
$Cr^{3+}_{(aq)}$, $Br^{-}_{(aq)}$, $H_2O_{(l)}$
SRA RA

$E°_r(\text{SOA}) = E°_r(\text{cathode})$. Likewise, $E°_r(\text{SRA}) = E°_r(\text{anode})$. To make the reaction begin, you must supply a potential that will at least counteract the negative cell potential of the electrolytic cell. That is why you take the additive inverse of the net cell potential. To calculate the cell potential, find out how far apart the oxidizing and reducing agents are.

SOA is $Cr^{3+}_{(aq)}$. $E°_r\left(Cr^{3+}_{(aq)}\right) = -0.41\text{ V}$

SRA is $Br^{-}_{(aq)}$. $E°_r\left(Br^{-}_{(aq)}\right) = +1.07\text{ V}$

$$\begin{aligned}\Delta E° &= E°_r(\text{SOA}) - E°_r(\text{SRA}) \\ &= -0.41\text{ V} - (+1.07\text{ V}) \\ &= -1.48\text{ V}\end{aligned}$$

Therefore, the minimum potential difference needed is +1.48 V.

b. SOA OA OA
$Cu_{(s)}$, $Cu^{2+}_{(aq)}$, $SO_4^{2-}{}_{(aq)}$, $H_2O_{(l)}$
SRA RA

SOA is $Cu^{2+}_{(aq)}$. $E°_r\left(Cu^{2+}_{(aq)}\right) = +0.34\text{ V}$

SRA is $Cu_{(s)}$. $E°_r\left(Cu_{(s)}\right) = +0.34\text{ V}$

$$\begin{aligned}\Delta E° &= E°_r(\text{SOA}) - E°_r(\text{SRA}) \\ &= +0.34\text{ V} - (+0.34\text{ V}) \\ &= 0.00\text{ V}\end{aligned}$$

Therefore, the minimum potential needed is 0.00 V. This means that any applied voltage greater than zero will make the cell run.

12. When a Ni-Cad battery runs a toy or other device, it is operating as a voltaic cell. The reactions involved are

Cathode: $2\,NiO(OH)_{(s)} + 2\,H_2O_{(l)} + 2\,e^- \rightarrow 2\,Ni(OH)_{2(s)} + 2\,OH^-_{(aq)}$

Anode: $Cd_{(s)} + 2\,OH^-_{(aq)} \rightarrow Cd(OH)_{2(s)} + 2\,e^-$

Net Reaction: $2\,NiO(OH)_{(s)} + 2\,H_2O_{(l)} + Cd_{(s)} \rightarrow 2\,Ni(OH)_{2(s)} + Cd(OH)_{2(s)}$

The net cell potential of this reaction is +1.25 V. As the battery operates, chemical energy is changed to electrical energy.

Eventually, the voltaic cell will not operate because the cell reactants are depleted to the point when a current can no longer be generated. At this point, the battery is considered dead, so it is plugged into a recharger. In the recharger, the electrical energy added into the battery gets converted back into chemical energy. In other words, the reaction is reversed. The cathode now becomes the anode and the anode now becomes the cathode. However, reduction is still occurring at the cathode and oxidation is still occurring at the anode. These reverse reactions replenish the original cell reactants. This is shown as follows:

Anode: $2\,Ni(OH)_{2(s)} + 2\,OH^-_{(aq)} \rightarrow 2\,NiO(OH)_{(s)} + 2\,H_2O_{(l)} + 2\,e^-$

Cathode: $Cd(OH)_{2(s)} + 2\,e^- \rightarrow Cd_{(s)} + 2\,OH^-_{(aq)}$

Net Reaction: $2\,Ni(OH)_{2(s)} + Cd(OH)_{2(s)} \rightarrow 2\,NiO(OH)_{(s)} + 2\,H_2O_{(l)} + Cd_{(s)}$

In effect, what you do when you recharge a battery is cause electrons to flow in the reverse from their spontaneous direction.

Section 2: Activity 2

1. a. The following are the predictions for the $CuSO_{4(aq)}$ cell.

SOA OA OA

$Cu^{2+}_{(aq)}$, $SO_4^{2-}_{(aq)}$, $H_2O_{(l)}$

SRA

Cathode (SOA): $2\left(Cu^{2+}_{(aq)} + 2\,e^- \rightarrow Cu_{(s)}\right)$ $E^\circ_r = +0.34$ V

Anode (SRA): $2\,H_2O_{(l)} \rightarrow O_{2(g)} + 4\,H^+_{(aq)} + 4\,e^-$ $E^\circ_{ox} = -1.23$ V

Net Reaction: $2\,Cu^{2+}_{(aq)} + 2\,H_2O_{(l)} \rightarrow O_{2(g)} + 4\,H^+_{(aq)} + 2\,Cu_{(s)}$ $\Delta E^\circ = -0.89$ V

or, $\Delta E^\circ = E^\circ_r(\text{cathode}) - E^\circ_r(\text{anode})$

$= +0.34\text{ V} - (+1.23\text{ V})$

$= -0.89$ V The net cell potential is predicted to be −0.89 V.

The following are the predictions for the $Na_2SO_{4(aq)}$ and $NaCl_{(aq)}$ cells. The cell predictions are identical.

OA OA SOA
$Na^+_{(aq)}$, $SO_4^{2-}{}_{(aq)}$, $H_2O_{(l)}$ **and** $Na^+_{(aq)}$, $Cl^-_{(aq)}$, $H_2O_{(l)}$
SRA SOA
RA RA SRA

Cathode (SOA): $2\left(2\,H_2O_{(l)} + 2\,e^- \rightarrow H_{2(g)} + 2\,OH^-_{(aq)}\right)$ $E°_r = -0.83$ V

Anode (SRA): $2\,H_2O_{(l)} \rightarrow O_{2(g)} + 4\,H^+_{(aq)} + 4\,e^-$ $E°_{ox} = -1.23$ V

Net Reaction: $2\,H_2O_{(l)} \rightarrow 2\,H_{2(g)} + O_{2(g)} + \cancel{4\,OH^-_{(aq)}} + \cancel{4\,H^+_{(aq)}}$ $\Delta E° = -2.06$ V

Because $OH^-_{(aq)} + H^+_{(aq)} \rightarrow H_2O_{(l)}$, this gives a net reaction of $2\,H_2O_{(l)} \rightarrow 2\,H_{2(g)} + O_{2(g)}$

or, $\Delta E° = E°_r(\text{cathode}) - E°_r(\text{anode})$
$= -0.83\text{ V} - (+1.23\text{ V})$
$= -2.06$ V

b. Evidence for the $CuSO_{4(aq)}$ cell: At the cathode, you should observe the blue colour slowly disappearing from the solution if left for a long enough time, and a reddish brown solid forming on the cathode. Some bubbles would be observed at the anode; the solution around the anode tests as being acidic $\left(H^+_{(aq)}\right)$.

Evidence for the $Na_2SO_{4(aq)}$ and $NaCl_{(aq)}$ cells: You should observe more bubbles forming at the cathode than at the anode. You should also observe that red litmus turns blue at the cathode, meaning the solution is becoming basic. Likewise, you should see that blue litmus turns red at the anode, meaning the solution is becoming acidic.

2. Your evidence table should be similar to the following:

Cell	Cathode	Anode
$CuSO_{4(aq)}$	• before operation – blue litmus turned red • after operation – blue litmus turned red; red-brown solid formed on electrode	• before operation – blue litmus turned red • after operation – blue litmus turned red; bubbles formed around electrode
$Na_2SO_{4(aq)}$	• before operation – no change in litmus • after operation – red litmus turned blue; vigorous bubbling around electrode	• before operation – no change in litmus • after operation – blue litmus turned red; bubbling around electrode but not as vigorous as cathode
$NaCl_{(aq)}$	• before operation – no change in litmus • after operation – red litmus turned blue; vigorous bubbling around electrode	• before operation – no change in litmus • after operation – both blue and red litmus turned white; some bubbling around electrode; strong smell like bleach; solution had a slight yellow-green colour

3. Your procedure should be similar to the following:

   - Add about 2 mL (1/2 teaspoon) of $CuSO_4 \bullet 5H_2O_{(s)}$ to 25 mL of water in a 100 mL beaker.
   - Mix the solution. Add about 20 mL of the $CuSO_{4(aq)}$ to a U-tube.
   - Set up the electrolytic cell apparatus as shown in Investigation 13.5 of Section 2, Activity 1.
   - Set the universal adapter to 3 V.
   - Clamp the electrodes carefully so that the carbon ends are immersed in the solution.
   - Record your initial observations.
   - Plug in the adapter and observe for two minutes. Record your observations.
   - Rinse the equipment.
   - Repeat the procedure for $Na_2SO_{4(aq)}$ and $NaCl_{(aq)}$.

4. See the evidence table in the answer for Section 2: Activity 2, question 2.

5. Your analysis should be similar to the following:

   $CuSO_{4(aq)}$ Cell: The evidence of a red-brown solid could confirm the presence of copper at the cathode. The bubbles support the possibility of $O_{2(g)}$, but this is not confirmed unless the gas is collected and tested. The evidence of $H^+_{(aq)}$ at the anode is not necessarily confirmed since blue litmus also turned slightly red before the cell was operated.

   $Na_2SO_{4(aq)}$ Cell: The predictions seem to be confirmed. However, since the gases weren't collected, it is not possible to confirm the gas identities.

   $NaCl_{(aq)}$ Cell: The cathode half-reaction seemed to be confirmed, again without testing the gas produced. However, the anode reaction was not confirmed. The white-coloured litmus paper and the smell of bleach suggests a different substance than expected was formed. Further tests are warranted to discover the identity of the anode products.

6. The following improvements could be made:

   - Increase the concentration of the reactants to 1.0 mol/L so the predictions of net cell potentials could have been made with more certainty.
   - Use a better low voltage power supply so that the potential across the electrodes could have been controlled more precisely, thereby minimizing the chances of additional side reactions.
   - Use a Hoffmann apparatus which would allow for the collection of gases at the electrodes. These gases could then be measured and tested for identity.
   - Use pH paper to more closely monitor the acidity at the anode of the $CuSO_{4(aq)}$ cell.

7. The process of oxidation must have produced the $Cl_{2(g)}$, since oxidation occurs at the anode.

8. There were chloride ions $\left(Cl^-_{(aq)}\right)$ present in the solution $\left(NaCl_{(aq)}\right)$ from which the chlorine could have been produced.

9. The possible half-reaction equation is as follows:

$$2\,Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2\,e^-$$

↑

Electrons lost indicate oxidation.

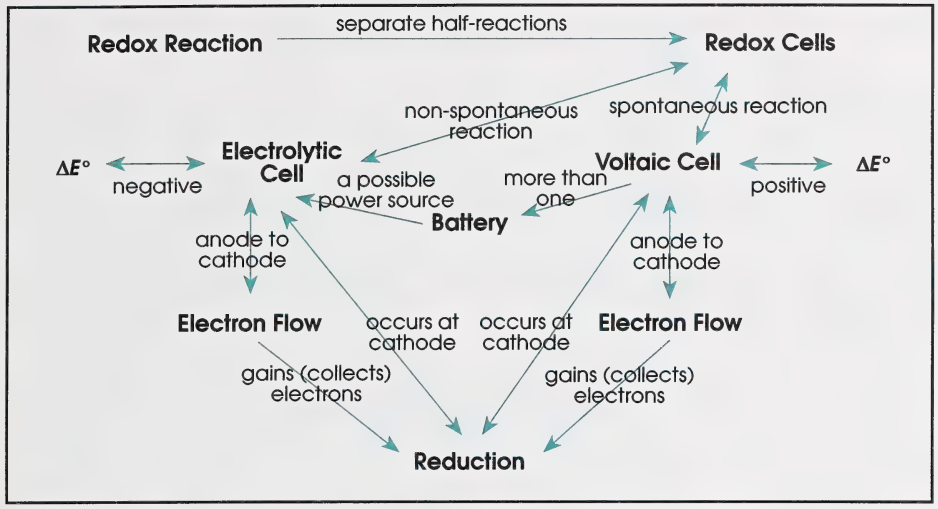
10. The chlorine half-reaction equation is only three places (and 0.13 V) above the water half-reaction equation. Therefore, water is not that much stronger than $Cl^-_{(aq)}$ as a reducing agent. Also, since the voltage of the power supply (3 V) was well above the minimum voltage required to run the cell ($-0.83\text{ V} - 1.36\text{ V} = -2.19\text{ V}$; therefore, minimum voltage = +2.19 V), it is possible that the chlorine gas could have been produced as follows:

Anode:	$2\,Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2\,e^-$	$E°_r = -1.36\text{ V}$
Cathode:	$2\,H_2O_{(l)} + 2\,e^- \rightarrow H_{2(g)} + 2\,OH^-_{(aq)}$	$E°_{ox} = -0.83\text{ V}$
Net Reaction:	$2\,Cl^-_{(aq)} + 2\,H_2O_{(l)} \rightarrow Cl_{2(g)} + H_{2(g)} + 2\,OH^-_{(aq)}$	$\Delta E° = -2.19\text{ V}$

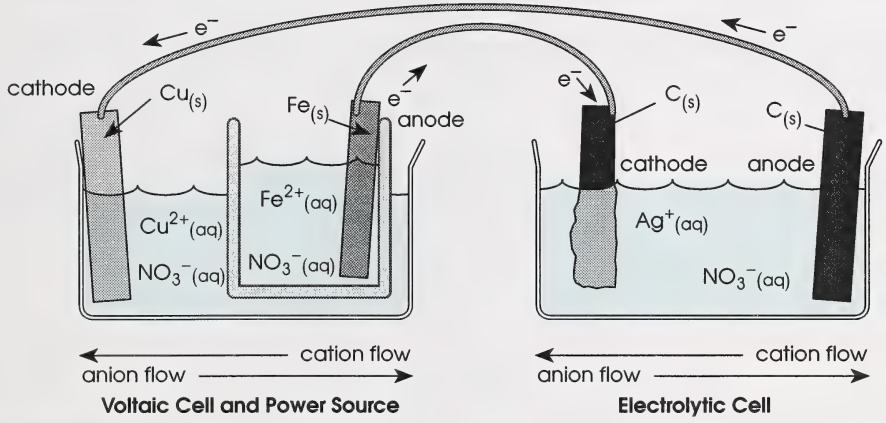
11.

Method of Metal Protection	Description	Advantages	Disadvantages
plating with less reactive metal	An unreactive metal coats another metal to make a barrier so materials cannot get in.	The unreactive metal will stay shiny.	It is very expensive.
plating with more reactive metal	A more reactive metal coats another metal to make a barrier so materials cannot get in.	It is fairly inexpensive.	The more reactive metal dulls as an oxide coating forms.
galvanizing	It involves coating the metal with zinc, then zinc forming its own protective coating.	It forms an inexpensive $ZnO_{(s)}$ coating.	It looks dull with an oxide coating.
cathodic protection	It involves connecting a more reactive metal to the protected metal.	The metal being protected is still visible; it is easy to maintain.	It works only for a localized area.

12.



13. a. Figure 4.4 should be labelled as follows:



b. To figure out whether electrons will actually flow, you need to consider the potential difference of the voltaic cell (left cell in Figure 4.4) versus the electrolytic cell (right cell in Figure 4.4).

Voltaic Cell

	SOA	OA	OA
$Cu_{(s)}$, $Fe_{(s)}$,	$Cu^{2+}_{(aq)}$,	$Fe^{2+}_{(aq)}$,	$NO_3^-{}_{(aq)}$, $H_2O_{(l)}$
RA	**SRA**	**RA**	**RA**

Cathode (SOA): $Cu^{2+}_{(aq)} + 2\,e^- \rightarrow Cu_{(s)}$ $\quad E^\circ_r = +0.34\ V$

Anode (SRA): $Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2\,e^-$ $\quad E^\circ_{ox} = +0.45\ V$

Net Reaction: $Cu^{2+}_{(aq)} + Fe_{(s)} \rightarrow Cu_{(s)} + Fe^{2+}_{(aq)}$ $\quad \Delta E^\circ = +0.79\ V$

or $\Delta E^\circ = +0.34\ V - (-0.45\ V)$
$= +0.79\ V$

Electrolytic Cell

SOA	OA
$Ag^+_{(aq)}$, $NO_3^-{}_{(aq)}$,	$H_2O_{(l)}$
	SRA

Cathode (SOA):	$4\left(Ag^+_{(aq)} + 1\,e^- \rightarrow Ag_{(s)}\right)$	$E°_r = +0.80$ V
Anode (SRA):	$2\,H_2O_{(l)} \rightarrow O_{2(g)} + 4\,H^+_{(aq)} + 4\,e^-$	$E°_{ox} = -1.23$ V
Net Reaction:	$4\,Ag^+_{(aq)} + 2\,H_2O_{(l)} \rightarrow 4\,Ag_{(s)} + O_{2(g)} + 4\,H^+_{(aq)}$	$\Delta E° = -0.43$ V

or, $\Delta E° = +0.80\text{ V} - (+1.23\text{ V})$
$= -0.43$ V

Minimum voltage required $= +0.43$ V

Since the voltaic cell in Figure 4.4 can provide a greater potential difference than the minimum required by the electrolytic cell, the electrons will flow to give a current.

14. The reactions for a discharging lead-acid cell are as follows:

Cathode:	$PbO_{2(s)} + 4\,H^+_{(aq)} + SO_4^{2-}{}_{(aq)} + 2\,e^- \rightarrow PbSO_{4(s)} + 2\,H_2O_{(l)}$	$E°_r = +1.69$ V
Anode:	$Pb_{(s)} + SO_4^{2-}{}_{(aq)} \rightarrow PbSO_{4(s)} + 2\,e^-$	$E°_{ox} = +0.36$ V
Net Reaction:	$PbO_{2(s)} + 4\,H^+_{(aq)} + 2\,SO_4^{2-}{}_{(aq)} + Pb_{(s)} \rightarrow 2\,PbSO_{4(s)} + 2\,H_2O_{(l)}$	$\Delta E° = +2.05$ V

or, $\Delta E° = E°_r(\text{cathode}) - E°_r(\text{anode})$
$= +1.69\text{ V} - (-0.36\text{ V})$
$= +2.05$ V

The cell potential is +2.05 V.

15. The reactions for a recharging lead-acid cell are as follows:

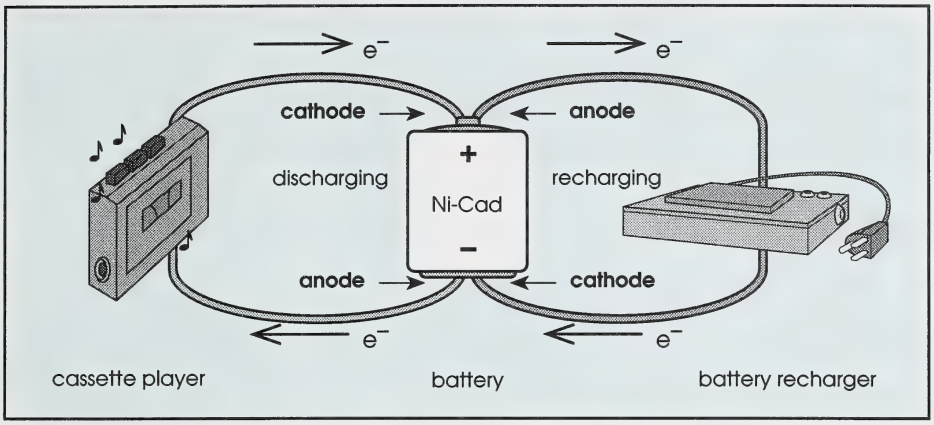
Cathode:	$PbSO_{4(s)} + 2\,e^- \rightarrow Pb_{(s)} + SO_4^{2-}{}_{(aq)}$	$E°_r = -0.36$ V
Anode:	$PbSO_{4(s)} + 2\,H_2O_{(l)} \rightarrow PbO_{2(s)} + 4\,H^+_{(aq)} + SO_4^{2-}{}_{(aq)} + 2\,e^-$	$E°_{ox} = -1.69$ V
Net Reaction:	$2\,PbSO_{4(s)} + 2\,H_2O_{(l)} \rightarrow PbO_{2(s)} + 4\,H^+_{(aq)} + Pb_{(s)} + 2\,SO_4^{2-}{}_{(aq)}$	$\Delta E° = -2.05$ V

or, $\Delta E° = E°_r(\text{cathode}) - E°_r(\text{anode})$
$= -0.36\text{ V} - (+1.69\text{ V})$
$= -2.05$ V

The minimum voltage required to recharge the battery is +2.05 V.

16. In comparing the two sets of reactions, you can see that both cases have the same participating substances, but that one set of reactions and the net potential is the reverse of the other set.

17.



Section 2: Activity 3

1. Answers will vary. The following are two possibilities. One thing in common between the apple juice and the aluminum foil is that they are both pure. One thing in common between the aluminum foil and the jewellery is that they are both metallic.

2. Cations always move to the **cathode**. Since cations are **positive** ions, when the cation of a binary ionic compound is reduced at the cathode, the element it forms is a **metal**. Anions, on the other hand, form **nonmetal** elements at the anode if they undergo **oxidation**.

3. The alternatives b, c, and d are all incorrect because the $E°$ values are for different phases of materials than those given in the redox table. For example, $Ni^{2+}_{(l)}$ is not the same as $Ni^{2+}_{(aq)}$, so it should not have the same $E°$ value. You may remember earlier being told that you could estimate $E°$ values sometimes for different states (electrolysis) but these were just estimates.

4. **Textbook question 30:**

   a. Two difficulties with producing metals from electrolysing aqueous solutions are

      - Water will often be a stronger oxidizing agent than the metal ions, so it will electrolyse first.
      - Many ionic substances are not soluble in water.

   b. To offset these difficulties, you can electrolyse the molten ionic compound, add a solvent (other than water) that will not be broken down first by electrolysis, or increase the potential difference so much that the metal will be produced anyway.

Textbook question 31:

The only species present are $Sc^{3+}_{(l)}$ and $Cl^{-}_{(l)}$.

Cathode: $2\left(Sc^{3+}_{(l)} + 3\,e^{-} \rightarrow Sc_{(l)}\right)$

Anode: $3\left(2\,Cl^{-}_{(l)} \rightarrow Cl_{2(g)} + 2\,e^{-}\right)$

Net Reaction: $2\,Sc^{3+}_{(l)} + 6\,Cl^{-}_{(l)} \rightarrow 2\,Sc_{(l)} + 3\,Cl_{2(g)}$

Textbook question 32:

Your answer should include points such as the following:

You should recycle metals because

- It costs a lot to obtain the pure metal from the raw ore.
- Metals are a finite resource in the world.
- Throwing away metals uses up more landfill space needed for other waste products.
- Purifying metals from the ores creates much more pollution.
- Recycling may reduce energy requirements (aluminum recycling uses only 5% of original energy).

However, on the negative side, recycling also eliminates jobs for Canadians who purify and extract the metal. Some recycling programs may cost more than producing a new product.

5. a. **Cathode:** $4\left(Ni^{3+}_{(l)} + 3\,e^{-} \rightarrow Ni_{(l)}\right)$

Anode: $3\left(2\,O^{2-}_{(l)} \rightarrow O_{2(g)} + 4\,e^{-}\right)$

Net Reaction: $4\,Ni^{3+}_{(l)} + 6\,O^{2-}_{(l)} \rightarrow 4\,Ni_{(l)} + 3\,O_{2(g)}$

b. **Cathode:** $2\left(Fe^{2+}_{(l)} + 2\,e^{-} \rightarrow Fe_{(l)}\right)$

Anode: $2\,O^{2-}_{(l)} \rightarrow O_{2(g)} + 4\,e^{-}$

Net Reaction: $2\,Fe^{2+}_{(l)} + 2\,O^{2-}_{(l)} \rightarrow 2\,Fe_{(l)} + O_{2(g)}$

c. **Cathode:** $2\left(Na^{+}_{(l)} + e^{-} \rightarrow Na_{(l)}\right)$

Anode: $2\,Cl^{-}_{(l)} \rightarrow Cl_{2(g)} + 2\,e^{-}$

Net Reaction: $2\,Na^{+}_{(l)} + 2\,Cl^{-}_{(l)} \rightarrow 2\,Na_{(l)} + Cl_{2(g)}$

d. **Cathode:** $2\left(Pb^{2+}_{(aq)} + 2\,e^{-} \rightarrow Pb_{(s)}\right)$

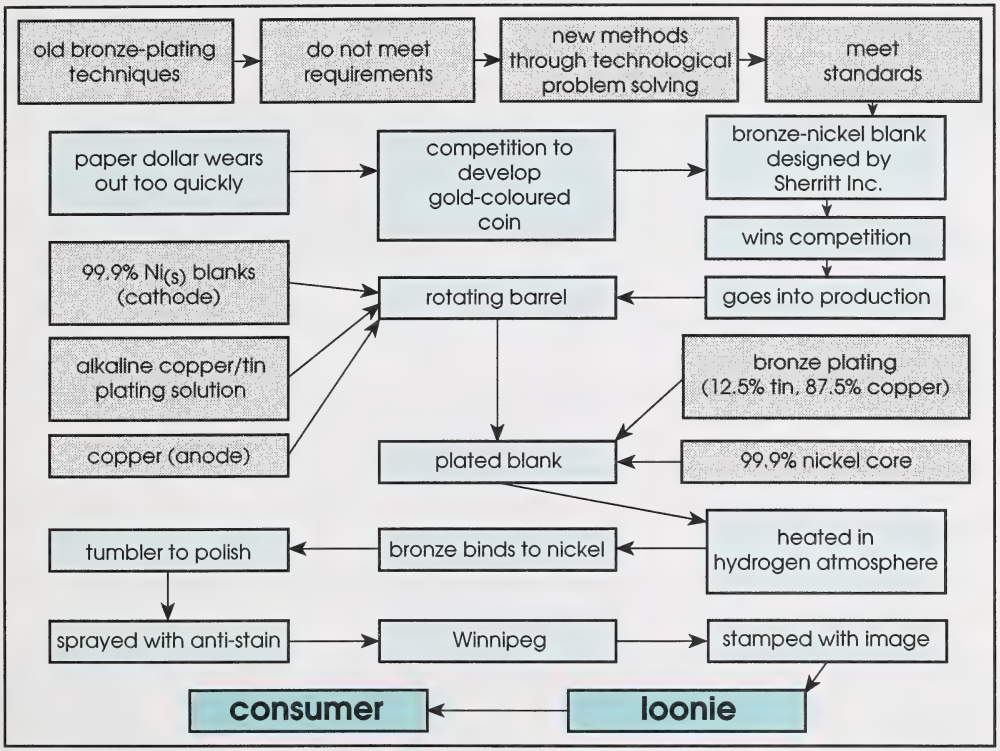
Anode: $2\,H_2O_{(l)} \rightarrow O_{2(g)} + 4\,H^{+}_{(aq)} + 4\,e^{-}$

Net Reaction: $2\,Pb^{2+}_{(aq)} + 2\,H_2O_{(l)} \rightarrow 2\,Pb_{(s)} + O_{2(g)} + 4\,H^{+}_{(aq)}$

6. a. iv b. v c. iii d. vi e. i f. ii

7. Only d. is false. The rest are true. Alternative d. is false because the metal to be electroplated should be at the cathode; that is where reduction occurs and where the cations will therefore go to be plated.

8.



9. Your prediction should take into account such things as concentration, voltage, and time. In addition, various additives such as acid that are put in the solution could help in this process. A sample set of variables is given for illustration. In each case the responding variable would be the degree of adhesion of the copper plating.

Trial	Manipulated Variable	Controlled Variables
1	voltage (3.0 V)	type of object (nail), time (0.5 h), apparatus (U-tube cell), volume of solution (20 mL), mass of $CuSO_4 \bullet 5H_2O$ (3 g)
2	voltage (4.5 V)	same controlled variables as Trial 1
3	voltage (6.0 V)	same controlled variables as Trial 1
4	concentration (0.10 mol/L)	same controlled variables minus concentration, plus voltage (3.0 V)
5	concentration (0.20 mol/L)	same controlled variables as Trial 4
6	concentration (0.50 mol/L)	same controlled variables as Trial 4

10. Materials lists will vary.

- possible objects to electroplate include
  - nails (uncoated, 5 cm or 2") } brass or stainless steel
  - washers (2 cm diameter) } brass or stainless steel
  - paper clips } brass or stainless steel
- steel wool
- $CuSO_4 \bullet 5H_2O_{(s)}$
- U-tube
- carbon electrode
- 2 leads
- universal adaptor
- retort stand
- 2 clamps
- 250 mL beaker

Your electrolytic cell could be a U-tube set up as in Investigations 13.5 or 13.6 (if you use nails or paper clips) or a beaker (if you use washers). Try to use only 20 mL of solution and between 2 mL and 5 mL (1/2-1 teaspoon) of $CuSO_4 \bullet 5H_2O_{(s)}$ for each trial.

11. Your procedure should be systematic, manipulating only one variable at a time, but each trial should include the following:

- Clean the metal object before electroplating.
- Attach your test metal as the cathode.
- Your anode should be carbon.
- Place the electrodes in the prepared solution.
- Plug in the universal adapter.
- Always electroplate for the same length of time unless time is a variable.
- Test the final product by rubbing it with steel wool.

12. Your evidence may be as follows:

A new washer was electroplated in each trial. Each washer was polished with steel wool and was electroplated for two minutes. The following data was recorded.

Potential Difference (Volts)	Concentration of $CuSO_{4(aq)}$ Solution			
	0.10 mol/L	0.20 mol/L	0.50 mol/L	0.50 mol/L (in 5% vinegar)
3.0	$Cu_{(s)}$ rubs off very easily.	$Cu_{(s)}$ rubs off easily.	$Cu_{(s)}$ rubs off easily.	$Cu_{(s)}$ rubs off with some difficulty.
4.5	$Cu_{(s)}$ rubs off easily.	$Cu_{(s)}$ rubs off moderately.	$Cu_{(s)}$ rubs off with some difficulty.	Some $Cu_{(s)}$ is still left after polishing.
6.0	$Cu_{(s)}$ rubs off easily.	$Cu_{(s)}$ rubs off moderately.	Some $Cu_{(s)}$ is still left after polishing.	The most $Cu_{(s)}$ is remaining out of all tests.

Note: For the last column, the $CuSO_4 \bullet 5H_2O_{(s)}$ was dissolved in household vinegar instead of water.

13. Based on the results, the best process for electroplating copper seems to be to use a 0.50 mol/L $CuSO_{4(aq)}$ solution containing 5% acetic acid (vinegar) with a six-volt potential difference.

14. Your evaluation may be similar to the following. While the method of determining the durability of the copper plating was crude, it was still effective. The way to improve the experiment would be to try electroplating the copper for various lengths of time, and at higher voltages in solutions at different acid strengths. This is because it appeared that the higher the voltage, the better the coating. Also, a greater acidity improved the coating. This path could be explored further.

Section 2: Activity 4

1. a.

River	Electrons
volume of river water that passes one point in one minute	**current**
how long you take to measure the river water volume; one day	**time**
total water volume moved in the river in one day	**total charge transferred**

b. The relationship between the three variables in the water relationship is as follows:

$$\text{total water volume in one day} = \frac{\text{volume of water}}{\text{minute}} \times \text{number of minutes (one day)}$$
$$= \text{river current} \times \text{time}$$

The relationship showing the electron relationship (total charge transferred) is as follows:

$$\text{total charge transferred} = \frac{\text{number of electrons}}{\text{second}} \times \text{time}$$
$$= \text{electric current} \times \text{time}$$

2. $$\text{Molar charge } (F) = \frac{1.6022 \times 10^{-19}\ \text{C}}{1\ \cancel{\text{electron}}} \times \frac{6.02 \times 10^{23}\ \cancel{\text{electrons}}}{1\ \text{mol}}$$
$$= 9.645\,244 \times 10^{4}\ \text{C/mol}$$
$$= 9.65 \times 10^{4}\ \text{C/mol}$$ The charge on one mole of electrons is 9.65×10^{4} C.

3. a. To find moles from mass and molar mass, you divide *mass* by *molar mass*.

$$\text{number of mol} = \frac{\text{number of g}}{\text{number of g/mol}} \quad \textbf{or} \quad \text{number of g} \times \frac{1\ \text{mol}}{\text{number of g}}$$

b. To find moles from heat and molar heat, you divide *heat* by *molar heat*.

$$\text{number of mol} = \frac{\text{number of kJ}}{\text{number of kJ/mol}} \quad \textbf{or} \quad \text{number of kJ} \times \frac{1\ \text{mol}}{\text{number of kJ}}$$

c. To find moles of electrons from charge and molar charge, you divide *charge* by *molar charge*.

$$\text{number of mol e}^{-}\ \left(n_{e^-}\right) = \frac{\text{number of C}}{\text{number of C/mol e}^{-}} \quad \textbf{or} \quad \text{number of C} \times \frac{1\ \text{mol e}^{-}}{\text{number of C}}$$

4. a. $n_{e^-} = 1.93\times10^4\ \cancel{C}\times\dfrac{1\text{ mol e}^-}{9.65\times10^4\ \cancel{C}}$

$= 0.200\text{ mol e}^-$

The number of moles of electrons is 0.200.

b. $n_{e^-} = 722\ \cancel{C}\times\dfrac{1\text{ mol e}^-}{9.65\times10^4\ \cancel{C}}$

$= 0.007\,481\,865\text{ mol e}^-$

$= 7.48\times10^{-3}\text{ mol e}^-$

The number of moles of electrons is 7.48×10^{-3}.

5. Another name for the charge on one mole of electrons is the **Faraday** constant. The letter used to represent this constant is **F**. It has units of **C/mol**. Likewise, current is measured in **amperes** and is represented by the letter **A**. One ampere is equivalent to one **coulomb** per second. The coulomb is a unit of charge that is represented by the letter **C**. When dealing with current, time is always measured in **seconds**.

6. a. $n_{e^-} = \dfrac{It}{F}$

$= \dfrac{(6.2\text{ C/s})(40\text{ s})}{9.65\times10^4\text{ C/mol}}$

$= 2.6\times10^{-3}\text{ mol e}^-$

or

$n_{e^-} = \dfrac{6.2\ \cancel{C}}{\cancel{s}}\times 40\ \cancel{s}\times\dfrac{1\text{ mol e}^-}{9.65\times10^4\ \cancel{C}}$

$= 0.002\,57\text{ mol e}^-$

$= 2.6\times10^{-3}\text{ mol e}^-$

b. $n_{e^-} = \dfrac{It}{F}$

$= \dfrac{(1.44\text{ C/s})\left(50\text{ min}\times\dfrac{60\text{ s}}{1\text{ min}}\right)}{9.65\times10^4\text{ C/mol}}$

$= 4.5\times10^{-2}\text{ mol e}^-$

or

$n_{e^-} = \dfrac{1.44\ \cancel{C}}{\cancel{s}}\times 50\ \cancel{\text{min}}\times\dfrac{60\ \cancel{s}}{1\ \cancel{\text{min}}}\times\dfrac{1\text{ mol e}^-}{9.65\times10^4\ \cancel{C}}$

$= 4.5\times10^{-2}\text{ mol e}^-$

c. $n_{e^-} = \dfrac{It}{F}$

$= \dfrac{\left(1.00\times10^3\text{ C/s}\right)\left(7\text{ h}\times\dfrac{60\text{ min}}{1\text{ h}}\times\dfrac{60\text{ s}}{1\text{ min}}\right)}{9.65\times10^4\text{ C/mol e}^-}$

$= 261.1\text{ mol e}^-$

$= 2.6\times10^2\text{ mol e}^-$

or

$n_{e^-} = \dfrac{1.00\times10^3\ \cancel{C}}{\cancel{s}}\times 7.0\ \cancel{h}\times\dfrac{60\ \cancel{\text{min}}}{1\ \cancel{h}}\times\dfrac{60\ \cancel{s}}{1\ \cancel{\text{min}}}\times\dfrac{1\text{ mol e}^-}{9.65\times10^4\ \cancel{C}}$

$= 261.1\text{ mol e}^-$

$= 2.6\times10^2\text{ mol e}^-$

7. **Textbook question 34:**

$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2\,e^-$

65.38 g/mol 0.500 A
m = ? 10.0 min
9.65×10^4 C/mol

$$m_{(Zn)} = \frac{0.500\ \cancel{C}}{\cancel{s}} \times 10.0\ \cancel{min} \times \frac{60\ \cancel{s}}{1\ \cancel{min}} \times \frac{1\ \cancel{mol\ e^-}}{9.65 \times 10^4\ \cancel{C}} \times \frac{1\ \cancel{mol\ Zn}}{2\ \cancel{mol\ e^-}} \times \frac{65.38\ g}{1\ \cancel{mol\ Zn}}$$
$$= 0.101\,626\,943\ g$$
$$= 0.102\ g$$

or

$$n_{e^-} = \frac{It}{F}$$
$$= \frac{(0.500\ C/s)\left(10\ min \times \frac{60\ s}{1\ min}\right)}{9.65 \times 10^4\ C/mol\ e^-}$$
$$= 3.1088 \times 10^{-3}\ mol\ e^-\ (rounded)$$

$$m_{(Cu)} = 3.1088 \times 10^{-3}\ \cancel{mol\ e^-} \times \frac{1\ \cancel{mol\ Zn}}{2\ \cancel{mol\ e^-}} \times \frac{65.38\ g}{1\ \cancel{mol\ Zn}}$$
$$= 0.102\ g$$

The mass of zinc oxidized was 0.102 g.

Textbook question 35:

a. $45\ min + 30\ s = (45\ min \times 60\ s/min) + 30\ s$
$= 2700\ s + 30\ s$
$= 2730\ s$

$Cr^{3+}_{(aq)} + 3\,e^- \rightarrow Cr_{(s)}$

54 A 52.00 g/mol
2730 s m = ?
9.65×10^4 C/mol

$$m_{(Cr)} = \frac{54\ \cancel{C}}{\cancel{s}} \times 2730\ \cancel{s} \times \frac{1\ \cancel{mol\ e^-}}{9.65 \times 10^4\ \cancel{C}} \times \frac{1\ \cancel{mol\ Cr}}{3\ \cancel{mol\ e^-}} \times \frac{52.00\ g}{1\ \cancel{mol\ Cr}}$$
$$= 26.48\ g$$
$$= 26\ g$$

or

$$n_{e^-} = \frac{It}{F}$$
$$= \frac{(54\ C/s)(2730\ s)}{9.65 \times 10^4\ C/mol}$$
$$= 1.528\ mol\ e^-\ (rounded)$$

$$m_{(Cr)} = 1.528\ \cancel{mol\ e^-} \times \frac{1\ \cancel{mol\ Cr}}{3\ \cancel{mol\ e^-}} \times \frac{52.00\ g}{1\ \cancel{mol\ Cr}}$$
$$= 26\ g$$

The mass of chromium deposited is 26 g.

b. $Ni^{2+}_{(aq)} + 2\,e^- \rightarrow Ni_{(s)}$

0.540 A	0.25 g
9.65×10^4 C/mol	58.69 g/mol
t = ?	

$$t = 0.25\ \cancel{g\ Ni} \times \frac{1\ \cancel{mol\ Ni}}{58.69\ \cancel{g\ Ni}} \times \frac{2\ \cancel{mol\ e^-}}{1\ \cancel{mol\ Ni}} \times \frac{9.65 \times 10^4\ \cancel{C}}{1\ \cancel{mol\ e^-}} \times \frac{1\ \cancel{s}}{0.540\ \cancel{C}} \times \frac{1\ min}{60\ \cancel{s}}$$

$= 25.37$ min

$= 25$ min

or

$$n_{e^-} = 0.25\ \cancel{g\ Ni} \times \frac{1\ \cancel{mol\ Ni}}{58.69\ \cancel{g\ Ni}} \times \frac{2\ mol\ e^-}{1\ \cancel{mol\ Ni}}$$

$= 0.008\ 519$ mol e^- (rounded)

$$n_{e^-} = \frac{It}{F}$$

$$t = \frac{n_{e^-} F}{I}$$

$$= \frac{(0.008\ 519\ mol\ e^-)(9.65 \times 10^4\ C/mol\ e^-)}{0.540\ C/s}$$

$= 1522\ \cancel{s} \times \frac{1\ min}{60\ \cancel{s}}$

$= 25$ min It took 25 min to plate the nickel.

c. $Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$

9.65×10^4 C/mol	10.00 g
84 min	107.87 g/mol
I = ?	

$$I = 10.00\ \cancel{g\ Ag} \times \frac{1\ \cancel{mol\ Ag}}{107.87\ \cancel{g\ Ag}} \times \frac{1\ \cancel{mol\ e^-}}{1\ \cancel{mol\ Ag}} \times \frac{9.65 \times 10^4\ C}{1\ \cancel{mol\ e^-}} \times \frac{1}{84\ \cancel{min}} \times \frac{1\ \cancel{min}}{60\ s}$$

$= 1.775$ C/s

$= 1.8$ A

or

$$n_{e^-} = 10.00\ \cancel{g\ Ag} \times \frac{1\ \cancel{mol\ Ag}}{107.87\ \cancel{g\ Ag}} \times \frac{1\ mol\ e^-}{1\ \cancel{mol\ Ag}}$$

$= 0.092\ 7042$ mol e^- (rounded)

$$n_{e^-} = \frac{It}{F}$$

$$I = \frac{n_{e^-} F}{t}$$

$$= \frac{(0.092\ 7042\ mol\ e^-)(9.65 \times 10^4\ C/mol\ e^-)}{84\ min \times \frac{60\ s}{1\ min}}$$

$= 1.8$ C/s

$= 1.8$ A The current required is 1.8 A.

Textbook question 36:

a. $Pb^{2+}_{(aq)} + 2\,e^- \rightarrow Pb_{(s)}$

1.00 A — 207.20 g/mol
120 h — m = ?
9.65×10^4 C/mol

$$m_{(Pb)} = \frac{1.00\ \cancel{C}}{\cancel{s}} \times 120\ \cancel{h} \times \frac{60\ \cancel{min}}{1\ \cancel{h}} \times \frac{60\ \cancel{s}}{1\ \cancel{min}} \times \frac{1\ \cancel{mol\ e^-}}{9.65 \times 10^4\ \cancel{C}} \times \frac{1\ \cancel{mol\ Pb}}{2\ \cancel{mol\ e^-}} \times \frac{207.20\ g}{1\ \cancel{mol\ Pb}}$$
$$= 463.78\ g$$
$$= 464\ g$$

or

$$n_{e^-} = \frac{It}{F}$$
$$= \frac{(1.00\ C/s)\left(120\ h \times \frac{3600\ s}{1\ h}\right)}{9.65 \times 10^4\ C/mol\ e^-}$$
$$= 4.4767\ mol\ e^-\ (rounded)$$

$$m_{(Pb)} = 4.4767\ \cancel{mol\ e^-} \times \frac{1\ mol\ Pb}{2\ \cancel{mol\ e^-}} \times \frac{207.20\ g}{1\ \cancel{mol\ Pb}}$$
$$= 464\ g$$

The mass of lead oxidized is 464 g.

b. $Al_{(s)} \rightarrow Al^{3+}_{(aq)} + 3\,e^-$

26.98 g/mol — 1.00 A
m = ? — 120 h
9.65×10^4 C/mol

$$m_{(Al)} = \frac{1.00\ \cancel{C}}{\cancel{s}} \times 120\ \cancel{h} \times \frac{60\ \cancel{min}}{1\ \cancel{h}} \times \frac{60\ \cancel{s}}{1\ \cancel{min}} \times \frac{1\ \cancel{mol\ e^-}}{9.65 \times 10^4\ \cancel{C}} \times \frac{1\ \cancel{mol\ Al}}{3\ \cancel{mol\ e^-}} \times \frac{26.98\ g}{1\ \cancel{mol\ Al}}$$
$$= 40.260\ g$$
$$= 40.3\ g$$

or

$$n_{e^-} = \frac{It}{F}$$
$$= \frac{(1.00\ C/s)\left(120\ h \times \frac{3600\ s}{1\ h}\right)}{9.65 \times 10^4\ C/mol}$$
$$= 4.4767\ mol\ e^-\ (rounded)$$

$$m_{(Al)} = 4.4767\ \cancel{mol\ e^-} \times \frac{1\ mol\ Al}{3\ \cancel{mol\ e^-}} \times \frac{26.98\ g}{1\ \cancel{mol\ Al}}$$
$$= 40.3\ g$$

The mass of aluminum oxidized is 40.3 g.

Textbook question 37:

In any redox reaction, the number of electrons lost must equal the number of electrons gained. Since this question does not involve electrical measurements, it becomes a simple stoichiometry question.

Cathode: $2\left(Al^{3+}_{(l)} + 3\,e^- \rightarrow Al_{(l)}\right)$

Anode: $3\left(2\,Cl^-_{(l)} \rightarrow Cl_{2(g)} + 2\,e^-\right)$

Net Reaction: $2\,Al^{3+}_{(l)} + 6\,Cl^-_{(l)} \rightarrow 2\,Al_{(l)} + 3\,Cl_{2(g)}$

	5.40 g	70.90 g/mol
	26.98 g/mol	m = ?

$$m_{(Cl_2)} = 5.40\ \cancel{g\ Al} \times \frac{1\ \cancel{mol\ Al}}{26.98\ \cancel{g\ Al}} \times \frac{3\ \cancel{mol\ Cl_2}}{2\ \cancel{mol\ Al}} \times \frac{70.90\ g}{1\ \cancel{mol\ Cl_2}}$$

$$= 21.286\ g$$

$$= 21.3\ g$$

The mass of chlorine produced at the anode is 21.3 g.

8. $Sn^{2+}_{(aq)} + 2\,e^- \rightarrow Sn_{(s)}$

3.46 A	118.69 g/mol
6.0 min	m = ?
9.65×10^4 C/mol	

$$m_{(Sn)} = \frac{3.46\ \cancel{C}}{\cancel{s}} \times 6.0\ \cancel{min} \times \frac{60\ \cancel{s}}{1\ \cancel{min}} \times \frac{1\ \cancel{mol\ e^-}}{9.65 \times 10^4\ \cancel{C}} \times \frac{1\ \cancel{mol\ Sn}}{2\ \cancel{mol\ e^-}} \times \frac{118.69\ g}{1\ \cancel{mol\ Sn}}$$

$$= 0.7660\ g$$

$$= 0.77\ g$$

or

$$n_{e^-} = \frac{It}{F}$$

$$= \frac{3.46\ C/\cancel{s}\left(6.0\ \cancel{min} \times \frac{60\ \cancel{s}}{1\ \cancel{min}}\right)}{9.65 \times 10^4\ C/mol}$$

$$= 0.012\ 91\ mol\ e^-\ \text{(rounded)}$$

$$m_{(Sn)} = 0.012\ 91\ \cancel{mol\ e^-} \times \frac{1\ \cancel{mol\ Sn}}{2\ \cancel{mol\ e^-}} \times \frac{118.69\ g}{1\ \cancel{mol\ Sn}}$$

$$= 0.77\ g$$

The predicted mass of tin plated at the cathode is 0.77 g.

9. $m_{(Sn)} = m_{final} - m_{initial}$

$= 118.05\ g - 117.34\ g$

$= 0.71\ g$ The actual mass of tin plated is 0.71 g.

10. a. $$\%\ \text{error} = \frac{|\text{experimental mass}| - |\text{theoretical mass}|}{|\text{theoretical mass}|} \times 100\%$$

$$= \frac{|0.71\ \text{g} - 0.77\ \text{g}|}{|0.77\ \text{g}|} \times 100\%$$

$= 7.8\%$ The percent error in the results is 7.8%.

b. A 7.8% error is a reasonable experimental error because of a number of factors. One factor that may cause less tin to plate out than theoretically predicted is that during the electrolysis, water may also have been electrolysed due to the size of the current. In addition, some of the tin plated on the steel can may not have adhered to the can. This tin could have been dislodged when the can was removed from the solution.

11. **$Au(NO_3)_3$ Solution**

$$1\left(Au^{3+}_{(aq)} + 3\,e^- \rightarrow Au_{(s)}\right)$$

$$Au^{3+}_{(aq)} + 3\,e^- \rightarrow Au_{(s)}$$

0.300 g

196.97 g/mol

$AgNO_3$ Solution

$$3\left(Ag^{+}_{(aq)} + e^- \rightarrow Ag_{(s)}\right)$$

$$3\,Ag^{+}_{(aq)} + 3\,e^- \rightarrow 3\,Ag_{(s)}$$

107.87 g/mol

m = ?

$$m_{(Ag)} = 0.300\ \cancel{\text{g Au}} \times \frac{1\ \cancel{\text{mol Au}}}{196.97\ \cancel{\text{g Au}}} \times \frac{3\ \cancel{\text{mol Ag}}}{1\ \cancel{\text{mol Au}}} \times \frac{107.87\ \text{g}}{1\ \cancel{\text{mol Ag}}}$$

$= 0.492\,88$ g

$= 0.493$ g

The mass of silver that can be plated is 0.493 g.

12. **Step 1:** As in question 11, the number of electrons transferred must be the same, so balance the electrons of the half-reactions as shown in the example and question 11.

$$6\left(Zn^{2+}_{(aq)} + 2\,e^- \rightarrow Zn_{(s)}\right) = 6\,Zn^{2+}_{(aq)} + 12\,e^- \rightarrow 6\,Zn_{(s)}$$

65.38 g/mol

m = ?

$$4\left(Ni^{3+}_{(aq)} + 3\,e^- \rightarrow Ni_{(s)}\right) = 4\,Ni^{3+}_{(aq)} + 12\,e^- \rightarrow 4\,Ni_{(s)}$$

58.69 g/mol

m = ?

$$3\left(Pb^{4+}_{(aq)} + 4\,e^- \rightarrow Pb_{(s)}\right) = 3\,Pb^{4+}_{(aq)} + 12\,e^- \rightarrow 3\,Pb_{(s)}$$

207.20 g/mol

m = ?

Step 2: Find out which has the greatest mass – 6 mol $Zn_{(s)}$, 4 mol $Ni_{(s)}$, or 3 mol $Pb_{(s)}$.

$$m_{(Zn)} = 6\ \cancel{\text{mol Zn}} \times \frac{65.38\ \text{g}}{1\ \cancel{\text{mol Zn}}} = 392.28\ \text{g}$$

$$m_{(Ni)} = 4\ \cancel{\text{mol Ni}} \times \frac{58.69\ \text{g}}{1\ \cancel{\text{mol Ni}}} = 234.76\ \text{g}$$

$$m_{(Pb)} = 3\ \cancel{\text{mol Pb}} \times \frac{207.20\ \text{g}}{1\ \cancel{\text{mol Pb}}} = 621.60\ \text{g}$$

The greatest mass produced will be the cell containing lead at 621.60 g.

13. The following explanations expand on the answers given in *Nelson Chemistry* (page 517).

Textbook question 24:

$Ni^{2+}_{(aq)}$	+	$2\ e^-$	→	$Ni_{(s)}$
0.500 L		3.50 A		
0.125 mol/L		9.65×10^4 C/mol		
		t = ?		

$$t = 0.500\ \cancel{\text{L Ni}^{2+}} \times \frac{0.125\ \cancel{\text{mol Ni}^{2+}}}{1\ \cancel{\text{L Ni}^{2+}}} \times \frac{2\ \cancel{\text{mol e}^-}}{1\ \cancel{\text{mol Ni}^{2+}}} \times \frac{9.65 \times 10^4\ \cancel{\text{C}}}{1\ \cancel{\text{mol e}^-}} \times \frac{1\ \cancel{\text{s}}}{3.50\ \cancel{\text{C}}} \times \frac{1\ \text{min}}{60\ \cancel{\text{s}}}$$

$$= 57.440\ \text{min}$$

$$= 57.4\ \text{min}$$

or

$$n_{e^-} = 0.500\ \cancel{\text{L Ni}^{2+}} \times \frac{0.125\ \cancel{\text{mol Ni}^{2+}}}{\cancel{\text{L Ni}^{2+}}} \times \frac{2\ \text{mol e}^-}{1\ \cancel{\text{mol Ni}^{2+}}}$$

$$= 0.125\ \text{mol e}^-$$

$$n_{e^-} = \frac{It}{F}$$

$$t = \frac{n_{e^-} F}{I}$$

$$= \frac{(0.125\ \text{mol e}^-)(9.65 \times 10^4\ \text{C/mol e}^-)}{3.50\ \text{C/s}}$$

$$= 3446\ \cancel{\text{s}} \times \frac{1\ \text{min}}{60\ \cancel{\text{s}}}$$

$$= 57.4\ \text{min}$$

It would take 57.4 min to plate the nickel.

Textbook question 25:

$Al^{3+}_{(cryolite)}$	+	$3\ e^-$	→	$Al_{(l)}$
		1.0 h		15 000 g
		9.65×10^4 C/mol		26.98 g/mol
		I = ?		

$$I = 15\,000\,\cancel{g} \times \frac{1\,\cancel{\text{mol Al}}}{26.98\,\cancel{g}} \times \frac{3\,\cancel{\text{mol e}^-}}{1\,\cancel{\text{mol Al}}} \times \frac{9.65 \times 10^4\ \text{C}}{1\,\cancel{\text{mol e}^-}} \times \frac{1}{1\,\cancel{\text{h}}} \times \frac{1\,\cancel{\text{h}}}{60\,\cancel{\text{min}}} \times \frac{1\,\cancel{\text{min}}}{60\ \text{s}}$$

$= 44\,709.043\,74\ \text{C/s (A)}$

$= 45\ \text{kA}$

or

$$n_{e^-} = 15\,000\,\cancel{\text{g Al}} \times \frac{1\,\cancel{\text{mol Al}}}{26.98\,\cancel{\text{g Al}}} \times \frac{3\ \text{mol e}^-}{1\,\cancel{\text{mol Al}}}$$

$= 1668\ \text{mol e}^-$ (rounded)

$$I = \frac{n_{e^-} F}{t}$$

$$= \frac{(1668\ \text{mol e}^-)(9.65 \times 10^4\ \text{C/mol e}^-)}{1.0\ \text{h} \times \frac{3600\ \text{s}}{\text{h}}}$$

$= 4.5 \times 10^4\ \text{C/s (A)}$

$= 45\ \text{kA}$

A 45 kA current would be required for 1 h to plate the aluminum.

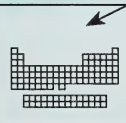
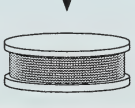
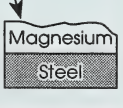
Section 2: Follow-up Activities

Extra Help

1.

a. B	d. B	g. V	j. E	m. B
b. B	e. B	h. V	k. E	n. B
c. B	f. V	i. E	l. B	o. E

2.

Cell	Electrolytic Cell (Electrolysis)			Voltaic Cell	
					
Use	element production	purification	electroplating	electric power source	cathodic protection
Example	chlorine	copper	Cr on steel	battery	Mg on Fe, or Zn on Fe

3. The order of the steps you follow to determine mass changes at the electrodes is as follows: c, b, e, a, d

4. $Zn^{2+}_{(aq)} + 2\,e^- \rightarrow Zn_{(s)}$

4.5 A	65.38 g/mol
20 min	m = ?
9.65×10^4 C/mol	

$$m_{(Zn)} = \frac{4.5\ \cancel{C}}{\cancel{s}} \times 20\ \cancel{min} \times \frac{60\ \cancel{s}}{1\ \cancel{min}} \times \frac{1\ \cancel{mol\ e^-}}{9.65 \times 10^4\ \cancel{C}} \times \frac{1\ \cancel{mol\ Zn}}{2\ \cancel{mol\ e^-}} \times \frac{65.38\ g}{1\ \cancel{mol\ Zn}}$$

$$= 1.829\ g$$

$$= 1.8\ g$$

or

$$n_{e^-} = \frac{It}{F}$$

$$= \frac{(4.5\ C/s)\left(20\ min \times \frac{60\ s}{1\ min}\right)}{9.65 \times 10^4\ C/mol}$$

$$= 0.055\ 96\ mol\ e^-\ (rounded)$$

$$m_{(Zn)} = 0.055\ 96\ \cancel{mol\ e^-} \times \frac{1\ \cancel{mol\ Zn}}{2\ \cancel{mol\ e^-}} \times \frac{65.38\ g}{1\ \cancel{mol\ Zn}}$$

$$= 1.8\ g$$

The mass of Zn gained by the cathode is 1.8 g.

5. $2\,O^{2-}_{(l)} \rightarrow O_{2(g)} + 4\,e^-$

32.00 g/mol	50.0 A
m = ?	1.5 h
	9.65×10^4 C/mol

$$m_{(O_2)} = \frac{50.0\ \cancel{C}}{\cancel{s}} \times 1.5\ \cancel{h} \times \frac{60\ \cancel{min}}{1\ \cancel{h}} \times \frac{60\ \cancel{s}}{1\ \cancel{min}} \times \frac{1\ \cancel{mol\ e^-}}{9.65 \times 10^4\ \cancel{C}} \times \frac{1\ \cancel{mol\ O_2}}{4\ \cancel{mol\ e^-}} \times \frac{32.00\ g}{1\ \cancel{mol\ O_2}}$$

$$= 22.38\ g$$

$$= 22\ g$$

or

$$n_{e^-} = \frac{It}{F}$$

$$= \frac{(50.0\ C/s)\left(1.5\ h \times \frac{3600\ s}{1\ h}\right)}{9.65 \times 10^4\ C/mol}$$

$$= 2.798\ mol\ e^-\ (rounded)$$

$$m_{(O_2)} = 2.798\ \cancel{mol\ e^-} \times \frac{1\ \cancel{mol\ O_2}}{4\ \cancel{mol\ e^-}} \times \frac{32.00\ g}{1\ \cancel{mol\ O_2}}$$

$$= 22\ g$$

The mass of $O_{2(g)}$ produced at the anode is 22 g.

Enrichment

1. Although Canada does not mine large quantities of aluminum, it imports and refines large quantities of a mineral called bauxite, the ore of aluminum, which contains about 33% aluminum by mass. Canada is able to do this because a tremendous amount of electricity is required for the process of purifying aluminum, and electricity is relatively cheap in Canada. In fact, it is estimated that approximately 60 MJ of energy are required to produce just one kilogram of aluminum once the aluminum oxide has been removed from the bauxite.

   The first step involved in extracting the aluminum oxide from the bauxite is to grind up the ore under pressure in the presence of sodium hydroxide, calcium oxide (lime), and sodium carbonate (soda ash). The compound sodium aluminate $\left(NaAlO_{2(aq)}\right)$ forms. The solution is then filtered to remove insoluble products.

   This same solution then undergoes a process called *hydrolysis* which causes $Al(OH)_{3(s)}$ to precipitate.

   Alumina $\left(Al_2O_{3(s)}\right)$ is produced from this product. Alumina is then electrolysed to form aluminum.

   The electrolysis process itself has only undergone minor changes since Charles Hall and Paul Héroult independently developed similar methods for refining aluminum in 1886. Before this, pure aluminum cost about \$300/kg (present dollar value). The reason for the high price was because of aluminum oxide's extremely high melting point (2015°C). By adding cryolite $\left(3\,NaF \bullet AlF_3\right)$ and other chemicals, the liquid mixture can now be electrolysed at around 960°C. With this reduction in temperature, the cost of aluminum ingots is just over \$2/kg.

   Recycling aluminum, however, requires only about five percent of the energy needed to extract aluminum from the bauxite ore. Recycling, therefore, results in a 95% energy savings over refining. This is predominantly because of an aluminum can's purity and low melting point of 660°C, which eliminates the large amount of pre-treatment and purification. To recycle aluminum, the aluminum cans are simply shipped to a smelter where they are melted down to remove slight impurities, then reformed and shipped back to manufacturers. Recycling also reduces air and water pollution, as well as conserving limited natural resources.

2. Except for copper used in the penny, the most common metal used in Canadian coins is nickel because it is relatively inexpensive and is quite durable. (Other countries, such as China, use iron for the centre of many of their coins.) However, all coins are made using the same three basic stages: purifying the metals, making the blanks, and finishing the coin.

   You can purify a metal in a number of ways. One is to electrolyse it by putting the impure metal at the anode, its metal ion as the electrolyte, and the pure metal at the cathode. Sherritt Inc. purifies its nickel by heating a nickel solution under high temperature and pressure in the presence of hydrogen. A redox reaction occurs that forms a nickel precipitate. After treating the nickel further, including heating it and rolling it to make it uniform, strips of nickel are formed from which the blanks of the coins can be made in various sizes. If the blanks need an extra coating of a different metal, as with the loonie, the process is also done at this time. In the case of the loonie, electrolysis is used to make a hard thin bronze coating. From here, the finished blanks are sent to the Canadian Mint in Winnipeg where the images are pressed into the coin and the coin edges are finished.

   As you can see, the only difference between making the loonie and the nickel is the fact that the loonie has to undergo additional steps to electroplate the bronze onto the nickel layer. Although electrolysis can be used to make both types of coins, it is not always the most economical way to manufacture things because of the high cost of electricity.

3. Your experiment should be similar to the following:

Investigation: Energy Comparisons of Batteries and Hydrocarbons

Problem

What mass of methane gas, when burned, will give the equivalent amount of energy as an AA-sized alkaline battery?

Prediction

It will take a relatively small mass of methane burned to release as much energy as an AA-sized alkaline battery.

Materials

- AA-sized alkaline battery*
- a plastic toy train*
- a watch or clock
- multimeter

*Any battery-operated toy with an appropriately sized battery may be used.

Procedure

- Record the voltage and the resistance of the battery using the multimeter.
- Place the battery in a small plastic toy train.
- Turn the train on and start recording the operating time.
- Stop recording the time when the toy no longer functions.
- Record the voltage and resistance of the battery again.
- Determine the average voltage and resistance.

Evidence

	Voltage (V)	Resistance (Ω)
Initial	1.462	1.1
Final	1.115	1.1
Average	1.289	1.1

Time elapsed = 7 h 43 min

Analysis

Using $E = V^2 t / R$, and the average readings, you can estimate the energy released by the battery.

$$E = \frac{(1.2885\ \text{V})^2\left(7.7167\ \text{h} \times \frac{60\ \text{min}}{\text{h}} \times \frac{60\ \text{s}}{\text{min}}\right)}{1.1\Omega}$$

$= 41\,928\ \text{J}$

$= 42\ \text{kJ}$ The energy released by the battery was 42 kJ.

The mass of methane releasing an equivalent amount of energy can be calculated as follows.

Calculate the molar heat of combustion of methane using Hess's law.

$$CH_{4(g)} + 2\,O_{2(g)} \rightarrow CO_{2(g)} + 2\,H_2O_{(l)}$$

$$\begin{aligned}\Delta H^\circ_c &= \sum nH^\circ_{fp} - \sum nH^\circ_{fr} \\ &= \left(2\ \cancel{\text{mol}} \times \frac{-285.8\ \text{kJ}}{1\ \cancel{\text{mol}}} + 1\ \cancel{\text{mol}} \times \frac{-393.5\ \text{kJ}}{1\ \cancel{\text{mol}}}\right) - \left(2\ \cancel{\text{mol}} \times \frac{0\ \text{kJ}}{1\ \cancel{\text{mol}}} + 1\ \cancel{\text{mol}} \times \frac{-74.4\ \text{kJ}}{1\ \cancel{\text{mol}}}\right) \\ &= -890.7\ \text{kJ} \\ &= -891\ \text{kJ}\end{aligned}$$

Since 891 kJ are released when 1 mole of methane burns, then $\Delta H^\circ_c = -891\ \text{kJ/mol}$.

To find the mass of CH_4 to give the equivalent of 42 kJ, you can use the equation $n = q / \Delta H$ and find mass from there, **or**

$$\begin{aligned}m_{(CH_4)} &= 41.928\ \cancel{\text{kJ}} \times \frac{1\ \cancel{\text{mol CH}_4}}{890.7\ \cancel{\text{kJ}}} \times \frac{16.05\ \text{g}}{1\ \cancel{\text{mol CH}_4}} \\ &= 0.76\ \text{g}\end{aligned}$$

Burning a 0.76 g mass of methane ($CH_{4(g)}$) will give the same amount of energy as the battery.

Evaluation

Since only an approximation was needed, this experiment worked well to show that a battery actually produces very little energy compared to hydrocarbon combustion. However, since the power loss is not linear, taking an average of initial and final readings was not accurate. Plotting a decay curve would have been more accurate. This was not necessary for this experiment.

NOTES

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